

Solutions

What You'll Learn

- ▶ You will describe and categorize solutions.
- ▶ You will calculate concentrations of solutions.
- ▶ You will analyze the colligative properties of solutions.
- ▶ You will compare and contrast heterogeneous mixtures.

Why It's Important

The air you breathe, the fluids in your body, and some of the foods you ingest are solutions. Because solutions are so common, learning about their behavior is fundamental to understanding chemistry.



Visit the Chemistry Web site at chemistrymc.com to find links about solutions.

Though it isn't apparent, there are at least three different solutions in this photo; the air, the lake in the foreground, and the steel used in the construction of the buildings are all solutions.



DISCOVERY LAB



Materials

balance
50-mL graduated cylinder
100-mL beaker (2)
stirring rod
ammonium chloride (NH_4Cl)
calcium chloride (CaCl_2)
water

Solution Formation

The intermolecular forces among dissolving particles and the attractive forces between solute and solvent particles result in an overall energy change. Can this change be observed?

Safety Precautions



Dispose of solutions by flushing them down a drain with excess water.

Procedure

1. Measure 10 g of ammonium chloride (NH_4Cl) and place it in a 100-mL beaker.
2. Add 30 mL of water to the NH_4Cl , stirring with your stirring rod.
3. Feel the bottom of the beaker and record your observations.
4. Repeat the procedure with calcium chloride (CaCl_2).

Analysis

Which dissolving process is exothermic? Endothermic? Suggest some practical applications for dissolving processes that are exothermic and for those that are endothermic.

Section

15.1

What are solutions?

Objectives

- **Describe** the characteristics of solutions and **identify** the various types.
- **Relate** intermolecular forces and the process of solvation.
- **Define** solubility and **identify** factors affecting it.

Vocabulary

soluble
insoluble
immiscible
miscible
solvation
heat of solution
solubility
saturated solution
unsaturated solution
supersaturated solution
Henry's law

Have you ever thought of the importance of solutions? Even if you haven't, the fact is that solutions are all around you. They are even inside you; you and all other organisms are composed of cells containing solutions that support life. You inhale a solution when you breathe. You are immersed in a solution whether standing in a room or swimming in a pool. Structures such as the ones shown in the photo on the previous page would not be possible without steel, yet another solution.

Characteristics of Solutions

Cell solutions, ocean water, and steel may appear quite dissimilar, but they share certain characteristics. In Chapter 3, you learned that solutions are homogeneous mixtures containing two or more substances called the solute and the solvent. The solute is the substance that dissolves. The solvent is the dissolving medium. When you look at a solution, it is not possible to distinguish the solute from the solvent.

A solution may exist as a gas, liquid, or solid depending on the state of its solvent, as shown in **Figure 15-1** and **Table 15-1** on the next page. Air is a gaseous solution, and its solvent is nitrogen gas; braces may be made of nitinol, a solid solution of titanium in nickel. Most solutions, however, are liquids. You learned in Chapter 10 that reactions can take place in aqueous solutions, that is, solutions in which reactants and products are mixed in water. In fact, water is the most common solvent among liquid solutions.

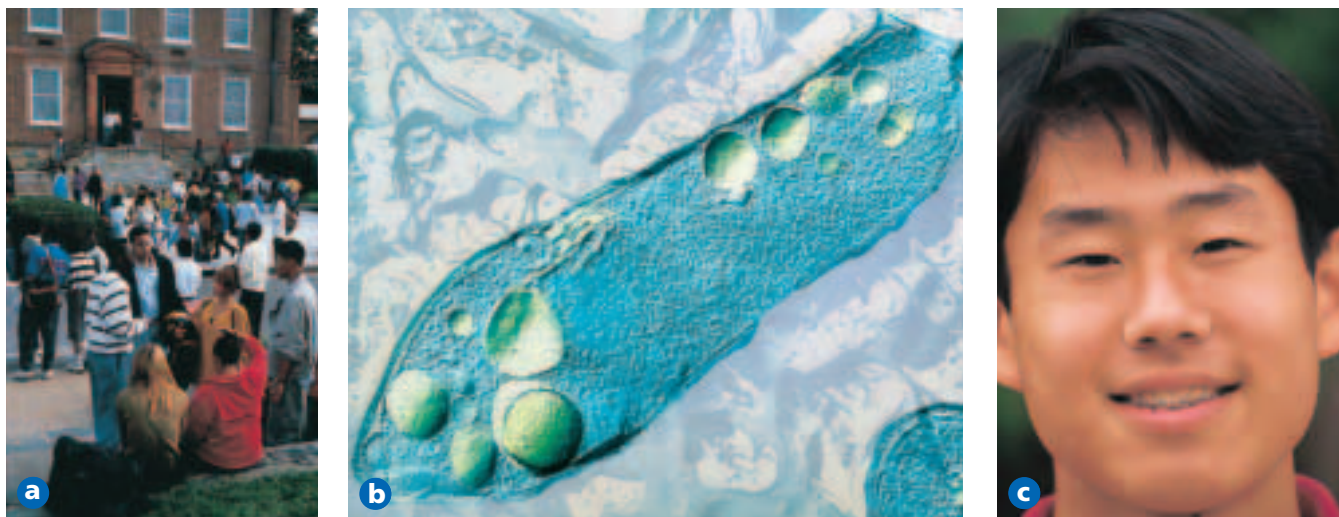


Figure 15-1

- a** The air you breathe is a gas solution primarily containing oxygen, nitrogen, and argon. **b** The biological reactions necessary for life occur in aqueous solutions within cells. **c** A solid solution of titanium and nickel is commonly used for braces in orthodontia.

Table 15-1 lists examples of different types of solutions. Note that the solutes in the solutions may be gases, liquids, or solids. Which solutions contain gaseous solutes? Which solutions are aqueous? Note also that solutions such as ocean water can contain more than one solute.

Some combinations of substances readily form solutions and others do not. A substance that dissolves in a solvent is said to be **soluble** in that solvent. For example, sugar is soluble in water, a fact you probably learned by dissolving sugar in flavored water to make a sweetened beverage such as tea, lemonade, or fruit punch. A substance that does not dissolve in a solvent is said to be **insoluble** in that solvent. Sand is insoluble in water. Have you ever shaken a bottle of oil and vinegar when making salad dressing? If so, what happens to the liquids shortly after you stop mixing them? You are correct if you answered that they separate, or cease to mix. Oil is insoluble in vinegar; and thus, oil and vinegar are said to be **immiscible**. Two liquids that are soluble in each other, such as those that form the antifreeze listed in **Table 15-1**, are said to be **miscible**. Are water and acetic acid miscible?

Table 15-1

Types and Examples of Solutions			
Type of solution	Example	Solvent	Solute
Gas			
Gas in gas	Air	Nitrogen (gas)	Oxygen (gas)
Liquid			
Gas in liquid	Carbonated water	Water (liquid)	Carbon dioxide (gas)
Gas in liquid	Ocean water	Water (liquid)	Oxygen gas (gas)
Liquid in liquid	Antifreeze	Water (liquid)	Ethylene glycol (liquid)
Liquid in liquid	Vinegar	Water (liquid)	Acetic acid (liquid)
Solid in liquid	Ocean water	Water (liquid)	Sodium chloride (solid)
Solid			
Liquid in solid	Dental amalgam	Silver (solid)	Mercury (liquid)
Solid in solid	Steel	Iron (solid)	Carbon (solid)

Solvation in Aqueous Solutions

Why are some substances soluble in one another whereas others are not? To form a solution, solute particles must separate from one another and the solute and solvent particles must mix. Recall from Chapter 13 that attractive forces exist among the particles of all substances. Attractive forces exist between the pure solute particles, between the pure solvent particles, and between the solute and solvent particles. When a solid solute is placed in a solvent, the solvent particles completely surround the surface of the solid solute. If the attractive forces between the solvent and solute particles are greater than the attractive forces holding the solute particles together, the solvent particles pull the solute particles apart and surround them. These surrounded solute particles then move away from the solid solute, out into the solution. The process of surrounding solute particles with solvent particles to form a solution is called **solvation**. Solvation in water is called hydration.

“Like dissolves like” is the general rule used to determine whether solvation will occur in a specific solvent. To determine whether a solvent and solute are alike, you must examine the bonding and the polarity of the particles and the intermolecular forces between particles.

Aqueous solutions of ionic compounds Examine **Figure 15-2** to review the polar nature of water. Recall that water molecules are dipoles with partially positive and partially negative ends. Water molecules are in constant motion as described by the kinetic-molecular theory. When a crystal of an ionic compound, such as sodium chloride (NaCl), is placed in a beaker of water, the water molecules collide with the surface of the crystal. The charged ends of the water molecules attract the positive sodium ions and negative chloride ions. This attraction between the dipoles and the ions is greater than the attraction among the ions in the crystal, and so the ions break away from the surface. The water molecules surround the ions and the solvated ions move into solution, as shown in **Figure 15-3**. This exposes more ions on the surface of the crystal. Solvation continues until the entire crystal has dissolved and all ions are distributed throughout the solvent.

Gypsum is a compound composed of calcium ions and sulfate ions. It is mixed with water to make plaster. Plaster is a mixture, not a solution. Gypsum is insoluble in water because the attractive forces among the ions in calcium sulfate are so strong that they cannot be overcome by the attractive forces exerted by the water molecules. Solvation does not occur.

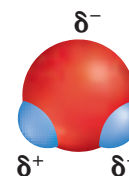


Figure 15-2

The bent shape of a water molecule results in dipoles that do not cancel each other out. The molecule has a net polarity, with the oxygen end being partially negative and the hydrogen ends being partially positive. Refer to **Table C-1** in Appendix C for a key to atom color conventions.

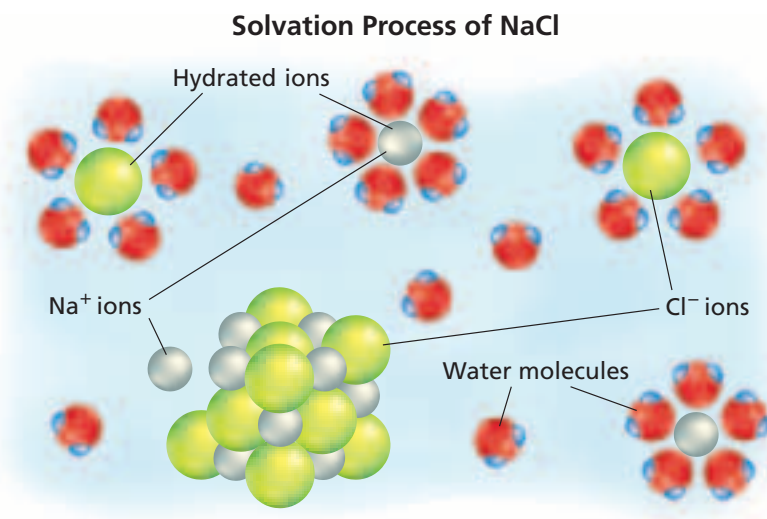


Figure 15-3

Solid sodium chloride dissolves as its ions are surrounded by solvent water molecules. Note how the polar water molecules orient themselves differently around the positive and negative ions.

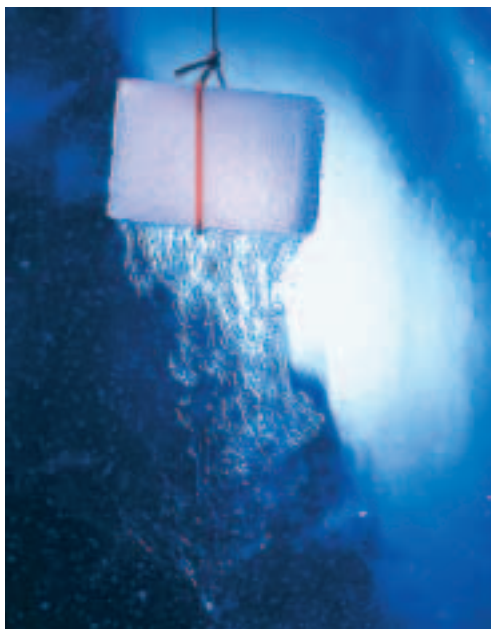
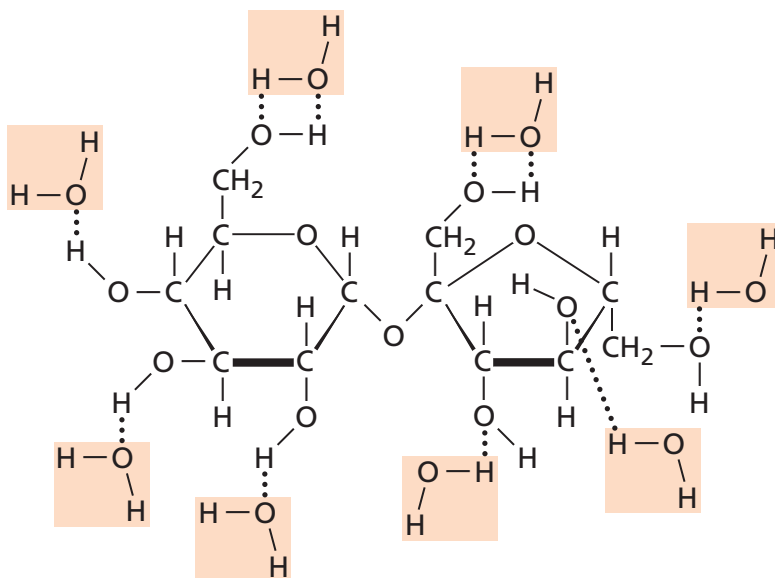


Figure 15-4

The solvation of a sugar cube (sucrose) can be seen in the above photo. Polar sucrose ($C_{12}H_{22}O_{11}$) molecules contain eight O–H bonds. Forces between polar water molecules and polar sucrose molecules break these O–H bonds and the sucrose dissolves.



Aqueous solutions of molecular compounds Water also is a good solvent for many molecular compounds. You know that table sugar dissolves in water. Table sugar is the molecular compound sucrose. See **Figure 15-4**. Note that its structure has a number of O–H bonds. Sucrose molecules are polar. When water is added to sucrose, each O–H bond becomes a site for hydrogen bonding with water. As soon as the sugar crystals contact the water, water molecules collide with the outer surface of the crystal. The attractive forces among sucrose molecules are overcome by the attractive forces between polar water molecules and polar sucrose molecules. Sugar molecules leave the crystal and become solvated by water molecules.

The oil in **Figure 15-5** is a substance made up of primarily carbon and hydrogen. It does not form a solution with water. Why are oil and water immiscible? There is little attraction between the polar water molecules and the nonpolar oil molecules. However, oil spills can be cleaned up with a nonpolar solvent. Nonpolar solutes are more readily dissolved in nonpolar solvents.



Figure 15-5

Nonpolar oil molecules do not mix with polar water molecules. Because oil is less dense than water, it floats on the water's surface. For this reason, oil spills at sea often wash ashore.

Factors that affect rate of solvation As you have just learned, solvation occurs only when and where the solute and solvent particles come in contact with each other. There are three common ways to increase the collisions between solute and solvent particles, and thus increase the rate at which the solute dissolves: agitating the mixture, increasing the surface area of the solute, and increasing the temperature of the solvent.

Agitating the mixture by stirring and shaking moves dissolved solute particles away from the contact surfaces more quickly and thereby allows new collisions between solute and solvent particles to occur. Without stirring or shaking, solvated particles move away from the contact areas slowly. Breaking the solute into small pieces increases its surface area. A greater surface area allows more collisions to occur. This is why a teaspoon of granulated sugar dissolves more quickly than an equal amount of sugar in a cube. Raising the temperature of the solvent increases the kinetic energy of its particles, resulting in more frequent collisions and collisions with greater energy than those that occur at lower temperatures. Can you think of an example of how an increase in temperature affects the rate of dissolving?

Heat of solution During the process of solvation, the solute must separate into particles. Solvent particles also must move apart in order to allow solute particles to come between them. Energy is required to overcome the attractive forces within the solute and within the solvent, so both steps are endothermic. When solute and solvent particles mix, the particles attract each other and energy is released. This step in the solvation process is exothermic. The overall energy change that occurs during the solution formation process is called the **heat of solution**.

As you observed in the **DISCOVERY LAB** at the beginning of this chapter, some solutions release energy as they form, whereas others absorb energy during formation. For example, after ammonium nitrate dissolves in water, its container feels cool. In contrast, after calcium chloride dissolves in water, its container feels warm. You will learn more about the heat of solution in the next chapter.

Solubility

If you have ever added so much sugar to a sweetened beverage that sugar crystals accumulated on the container's bottom, then you know that only a limited amount of solute can dissolve in a solvent at a given set of conditions. In fact, every solute has a characteristic solubility. **Solubility** refers to the maximum amount of solute that will dissolve in a given amount of solvent at a specified temperature and pressure. As you can see from **Table 15-2**, solubility is usually expressed in grams of solute per 100 g of solvent.

Table 15-2

Solubilities of Some Solutes in Water at Various Temperatures					
Substance	Formula	Solubility (g/100 g H ₂ O)*			
		0°C	20°C	60°C	100°C
Aluminum sulfate	Al ₂ (SO ₄) ₃	31.2	36.4	59.2	89.0
Ammonium chloride	NH ₄ Cl	29.4	37.2	55.3	77.3
Barium hydroxide	Ba(OH) ₂	1.67	3.89	20.94	—
Barium nitrate	Ba(NO ₃) ₂	4.95	9.02	20.4	34.4
Calcium hydroxide	Ca(OH) ₂	0.189	0.173	0.121	0.076
Lead(II) chloride	PbCl ₂	0.67	1.00	1.94	3.20
Lithium sulfate	Li ₂ SO ₄	36.1	34.8	32.6	—
Potassium chloride	KCl	28.0	34.2	45.8	56.3
Potassium sulfate	K ₂ SO ₄	7.4	11.1	18.2	24.1
Sodium chloride	NaCl	35.7	35.9	37.1	39.2
Silver nitrate	AgNO ₃	122	216	440	733
Sucrose	C ₁₂ H ₂₂ O ₁₁	179.2	203.9	287.3	487.2
Ammonia*	NH ₃	1130	680	200	—
Carbon dioxide*	CO ₂	1.713	0.878	0.359	—
Oxygen*	O ₂	0.048	0.031	0.019	—

* L/1 L H₂O of gas at standard pressure (101 kPa)

— No value available

Earth Science

CONNECTION

Water is the most abundant solvent on Earth. The amount of oxygen contained in water is called dissolved oxygen (DO). Dissolved oxygen is an indicator of water quality because oxygen is necessary for fish and other aquatic life. Oxygen enters a body of water by photosynthesis of aquatic plants and by transfer of oxygen across the air-water boundary.

Thermal pollution is a reduction in water quality due to an increase in water temperature. The discharge of heated water into a lake, river, or other body of water by factories or power plants causes the solubility of oxygen in water to decrease. In addition to not having enough oxygen to support life, as shown below, the lower oxygen level magnifies the effects of toxic and organic pollutants.

To reduce thermal pollution, measures must be taken before the heated water is released into the environment. The water can be discharged into holding lakes or canals and allowed to cool. In addition, many facilities use cooling towers to dissipate heat into the air.



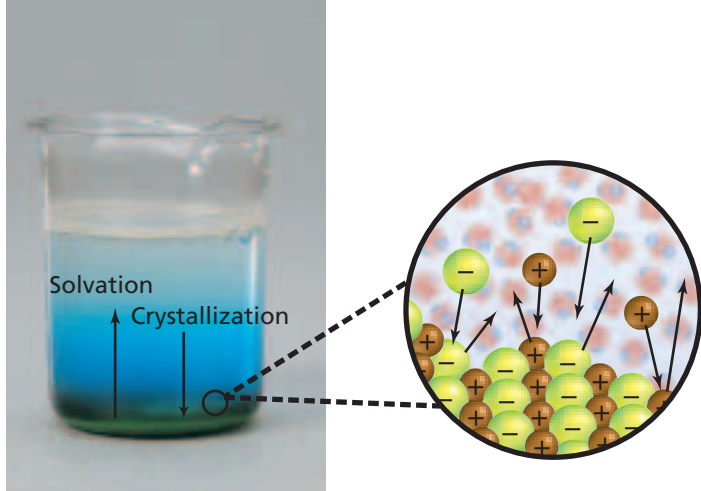


Figure 15-6

A dynamic equilibrium exists in a saturated solution. That is, the rate at which solute particles in the crystal are solvated is equal to the rate at which solvated solute particles rejoin the crystal.

Just as solvation can be understood at the particle level, so can solubility. When a solute is added to a solvent, solvent particles collide with the solute's surface particles; solute particles begin to mix randomly among the solvent particles. At first, the solute particles are carried away from the crystal. However, as the number of solvated particles increases, the same random mixing results in increasingly frequent collisions between solvated solute particles and the remaining crystal. Some colliding solute particles rejoin the crystal, or crystallize as you can see in **Figure 15-6**. As solvation continues, the crystallization rate increases while the solvation rate remains constant. As long as the solvation rate is greater than the crystallization rate, the net effect is continuing solvation.

Depending on the amount of solute present, the rates of solvation and crystallization may eventually equalize: no more solute appears to dissolve and a state of dynamic equilibrium exists between crystallization and solvation (as long as the temperature remains constant). Although solute particles continue to dissolve and crystallize in solutions that reach equilibrium, the overall amount of dissolved solute in the solution remains constant. Such a solution is said to be a **saturated solution**; it contains the maximum amount of dissolved solute for a given amount of solvent at a specific temperature and pressure. An **unsaturated solution** is one that contains less dissolved solute for a given temperature and pressure than a saturated solution. In other words, more solute can be dissolved in an unsaturated solution.

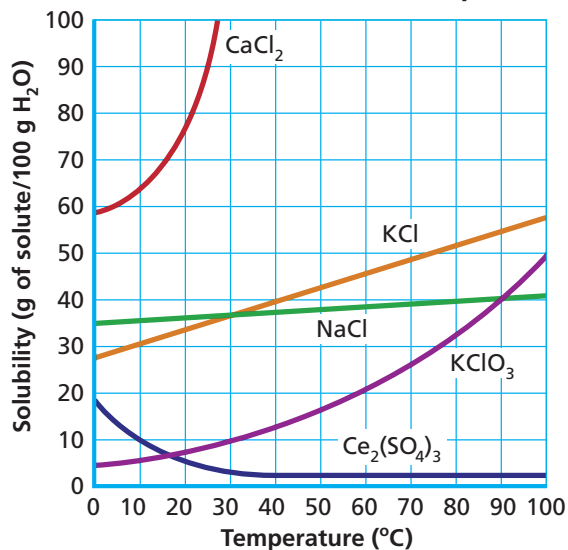
Factors That Affect Solubility

Figure 15-7

This graph shows the solubility of several substances as a function of temperature. What is the solubility of KClO_3 at 60°C ?

Pressure affects the solubility of gaseous solutes and gaseous solutions. The solubility of a solute also depends on the nature of the solute and solvent. Temperature affects the solubility of all substances. To learn how a blood disorder called sickle-cell disease can affect oxygen's solubility in blood, read the **Chemistry and Society** feature at the end of this chapter.

Solubilities as a Function of Temperature



Temperature and solubility Many substances are more soluble at high temperatures than at low temperatures, as you can see by the data graphed in **Figure 15-7**. For example, calcium chloride (CaCl_2) has a solubility of about 64 g CaCl_2 per 100 g H_2O at 10°C . Increasing the temperature to approximately 27°C increases the solubility by 50%, to 100 g CaCl_2 per 100 g H_2O . This fact also is illustrated by the data in **Table 15-2**, on the previous page. From the table you can see that at 20°C , 203.9 g of sucrose ($\text{C}_{12}\text{H}_{22}\text{O}_{11}$) dissolves in 100 g of water. At 100°C , 487.2 g of sucrose dissolves in 100 g of water, nearly a 140% increase in solubility over an 80°C temperature range. According to **Figure 15-7** and **Table 15-2**, what substances decrease in solubility as temperature increases?



Figure 15-8

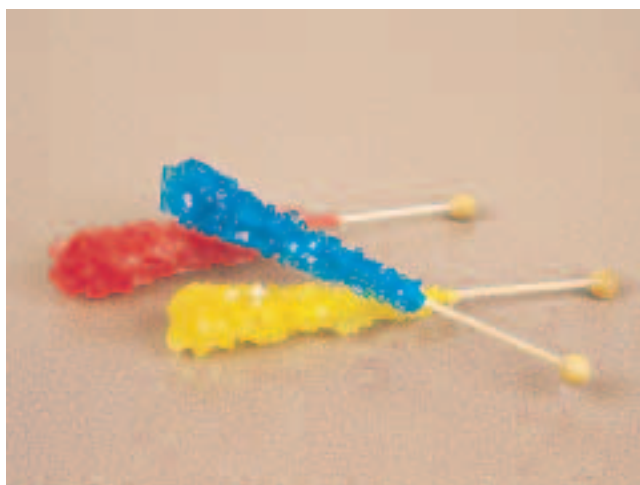
Sodium acetate ($\text{NaC}_2\text{H}_3\text{O}_2$) is commonly used in the preparation of a supersaturated solution. When a seed crystal is added **a**, the excess sodium acetate quickly crystallizes out of the solution, **b** and **c**.

As you can see from **Table 15-2**, the gases oxygen and carbon dioxide are less soluble at higher temperatures than at lower temperatures. This is a predictable trend for all gaseous solutes in liquid solvents. Can you explain why? Recall from Chapter 13 that the kinetic energy of gas particles allows them to escape from a solution more readily at higher temperatures. Thus, as a solution's temperature increases, the solubility of a gaseous solute decreases.

The fact that solubility changes with temperature and that some substances become more soluble with increasing temperature, is the key to forming supersaturated solutions. A **supersaturated solution** contains more dissolved solute than a saturated solution at the same temperature. To make a supersaturated solution, a saturated solution is formed at a high temperature and then cooled slowly. The slow cooling allows the excess solute to remain dissolved in solution at the lower temperature, as shown in **Figure 15-8a**. Supersaturated solutions are unstable. If a tiny amount of solute, called a seed crystal, is added to a supersaturated solution, the excess solute precipitates quickly, as shown in **Figures 15-8b** and **15-8c**. Crystallization can also occur if the inside of the container is scratched or the supersaturated solution undergoes a physical shock such as stirring or tapping the container. Using crystals of silver iodide (AgI) to seed air supersaturated with water vapor causes the water particles to come together and form droplets that may fall to Earth as rain. This often-performed technique is called cloud seeding. The rock candy and mineral deposits at the edges of mineral springs, shown in **Figure 15-9**, are formed from supersaturated solutions.

Figure 15-9

Examples of crystals that formed from supersaturated solutions include rock candy and hot spring mineral deposits.



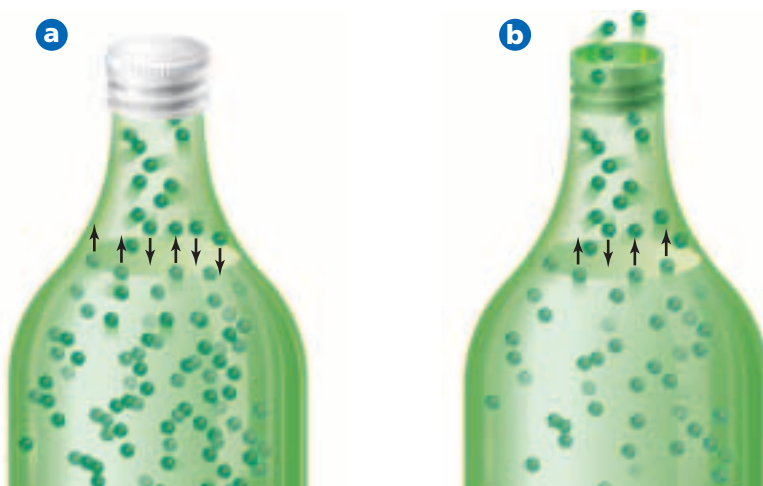


Figure 15-10

a When the cap on the soda bottle is closed, pressure above the solution keeps excess carbon dioxide (CO_2) from escaping the solution. **b** When the cap is removed, the decreased pressure above the solution results in the decreased solubility of the carbon dioxide—the carbon dioxide escapes the solution.

Pressure and solubility Pressure affects the solubility of gaseous solutes. The solubility of a gas in any solvent increases as its external pressure (the pressure above the solution) increases. Carbonated beverages depend on this fact. Carbonated beverages contain carbon dioxide gas dissolved in an aqueous solution. The dissolved gas gives the beverage its fizz. In bottling the beverage, carbon dioxide is dissolved in the solution at a pressure higher than atmospheric pressure. When the beverage container is opened, the pressure of the carbon dioxide gas in the space above the liquid (in the neck of the bottle) decreases. As a result, bubbles of carbon dioxide gas form in the solution, rise to the top, and escape. See **Figure 15-10**. Unless the cap is placed back on the bottle, the process will continue until the solution loses almost all of its carbon dioxide gas and goes flat.

Henry's law The decreased solubility of the carbon dioxide contained in the beverage after its cap is removed can be described by Henry's law. **Henry's law** states that at a given temperature, the solubility (S) of a gas in a liquid is directly proportional to the pressure (P) of the gas above the liquid. You can express this relationship in the following way

$$\frac{S_1}{P_1} = \frac{S_2}{P_2}$$

where S_1 is the solubility of a gas at a pressure P_1 and S_2 is the solubility of the gas at the new pressure P_2 .

You often will solve Henry's law for the solubility S_2 at a new pressure P_2 , where P_2 is known. The basic rules of algebra can be used to solve Henry's law for any one specific variable. To solve for S_2 , begin with the standard form of Henry's law.

$$\frac{S_1}{P_1} = \frac{S_2}{P_2}$$

Cross-multiplying yields,

$$S_1 P_2 = P_1 S_2$$

Dividing both sides of the equation by P_1 yields the desired result, the equation solved for S_2 .

$$\frac{S_1 P_2}{P_1} = \frac{P_1 S_2}{P_1} \quad S_2 = \frac{S_1 P_2}{P_1}$$



Topic: Solutions in the Body

To learn more about solutions in the body, visit the Chemistry Web site at chemistrymc.com

Activity: Research the solutions that can be found in the human body and in animals. How do biological systems make use of the principles of solubility to support life?

EXAMPLE PROBLEM 15-1

Using Henry's Law

If 0.85 g of a gas at 4.0 atm of pressure dissolves in 1.0 L of water at 25°C, how much will dissolve in 1.0 L of water at 1.0 atm of pressure and the same temperature?

1. Analyze the Problem

You are given the solubility of a gas at an initial pressure. The temperature of the gas remains constant as the pressure changes. Because decreasing pressure reduces a gas's solubility, less gas should dissolve at the lower pressure.

Known

$$S_1 = 0.85 \text{ g/L}$$

$$P_1 = 4.0 \text{ atm}$$

$$P_2 = 1.0 \text{ atm}$$

Unknown

$$S_2 = ? \text{ g/L}$$

2. Solve for the Unknown

Rearrange Henry's law to solve for S_2 .

$$\frac{S_1}{P_1} = \frac{S_2}{P_2}, \quad S_2 = S_1 \left(\frac{P_2}{P_1} \right)$$

Substitute the known values into the equation and solve.

$$S_2 = (0.85 \text{ g/L}) \left(\frac{1.0 \text{ atm}}{4.0 \text{ atm}} \right) = 0.21 \text{ g/L}$$

3. Evaluate the Answer

The answer is correctly expressed to two significant figures. The solubility decreased as expected. The pressure on the solution was reduced from 4.0 atm to 1.0 atm, so the solubility should be reduced to one-fourth its original value, which it is.

PRACTICE PROBLEMS

1. If 0.55 g of a gas dissolves in 1.0 L of water at 20.0 kPa of pressure, how much will dissolve at 110.0 kPa of pressure?
2. A gas has a solubility of 0.66 g/L at 10.0 atm of pressure. What is the pressure on a 1.0-L sample that contains 1.5 g of gas?



For more practice with Henry's law problems, go to **Supplemental Practice Problems** in Appendix A.

Section 15.1 Assessment

3. Describe the characteristics of a solution and identify the various types.
4. How do intermolecular forces affect solvation?
5. What is solubility? Describe two factors that affect solubility.
6. **Thinking Critically** If a seed crystal was added to a supersaturated solution, how would you characterize the resulting solution?

7. **Making and Using Graphs** Use the information in **Table 15-2** to graph the solubilities of aluminum sulfate, lithium sulfate, and potassium sulfate at 0°C, 20°C, 60°C, and 100°C. Which substance's solubility is most affected by increasing temperature?



Objectives

- **State** the concentrations of solutions in different ways.
- **Calculate** the concentrations of solutions.

Vocabulary

concentration
molarity
molality
mole fraction

You have learned about the process of solvation and the factors that affect solubility. The **concentration** of a solution is a measure of how much solute is dissolved in a specific amount of solvent or solution. How would you describe the concentration of the solutions in **Figure 15-11**? Concentration may be described qualitatively using the words *concentrated* or *dilute*. In general, a concentrated solution, as shown on the left in **Figure 15-11**, contains a large amount of solute. Conversely, a dilute solution contains a small amount of solute. How do you know that the tea on the right in **Figure 15-11** is a more dilute solution than the tea on the left?

Expressing Concentration

Although qualitative descriptions of concentration can be useful, solutions are more often described quantitatively. Some commonly used quantitative descriptions are percent by either mass or volume, molarity, and molality. These descriptions express concentration as a ratio of measured amounts of solute and solvent or solution. **Table 15-3** lists each ratio's description.

Table 15-3

Concentration Ratios	
Concentration description	Ratio
Percent by mass	$\frac{\text{mass of solute}}{\text{mass of solution}} \times 100$
Percent by volume	$\frac{\text{volume of solute}}{\text{volume of solution}} \times 100$
Molarity	$\frac{\text{moles of solute}}{\text{liter of solution}}$
Molality	$\frac{\text{moles of solute}}{\text{kilogram of solvent}}$
Mole fraction	$\frac{\text{moles of solute}}{\text{moles of solute} + \text{moles of solvent}}$

You may be wondering if one description is preferable to another. The description used depends on the type of solution analyzed and the reason for describing it. For example, a chemist working with a reaction in an aqueous solution most likely refers to the molarity of the solution, because he or she needs to know the number of particles involved in the reaction.

Figure 15-11

The strength of the tea corresponds to its concentration. The darker cup of tea is more concentrated than the lighter cup.



Using Percent to Describe Concentration

Concentration expressed as a percent is a ratio of a measured amount of solute to a measured amount of solution. Percent by mass usually describes solutions in which a solid is dissolved in a liquid, such as sodium chloride in water. The percent by mass is the ratio of the solute's mass to the solution's mass expressed as a percent. The mass of the solution equals the sum of the masses of the solute and the solvent.

$$\text{Percent by mass} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100$$

EXAMPLE PROBLEM 15-2

Calculating Percent by Mass

In order to maintain a sodium chloride (NaCl) concentration similar to ocean water, an aquarium must contain 3.6 g NaCl per 100.0 g of water. What is the percent by mass of NaCl in the solution?

1. Analyze the Problem

You are given the amount of sodium chloride dissolved in 100.0 g of water. The percent by mass of a solute is the ratio of the solute's mass to the solution's mass, which is the sum of the masses of the solute and the solvent.

Known

mass of solute = 3.6 g NaCl

mass of solvent = 100.0 g H₂O

Unknown

percent by mass = ?

2. Solve for the Unknown

Find the mass of the solution.

$$\begin{aligned}\text{Mass of solution} &= \text{grams of solute} + \text{grams of solvent} \\ &= 3.6 \text{ g} + 100.0 \text{ g} \\ &= 103.6 \text{ g}\end{aligned}$$

Substitute the known values into the percent by mass equation.

$$\begin{aligned}\text{Percent by mass} &= \frac{\text{mass of solute}}{\text{mass of solution}} \times 100 \\ &= \frac{3.6 \text{ g}}{103.6 \text{ g}} \times 100 \\ &= 3.5\%\end{aligned}$$

3. Evaluate the Answer

Because only a small mass of sodium chloride is dissolved per 100.0 g of water, the percent by mass should be a small value, which it is. The mass of sodium chloride was given with two significant figures, therefore, the answer also is expressed with two significant figures.



Maintaining the proper saline (salt) concentration is important to the health of saltwater fish.

PRACTICE PROBLEMS

- What is the percent by mass of NaHCO₃ in a solution containing 20 g NaHCO₃ dissolved in 600 mL H₂O?
- You have 1500.0 g of a bleach solution. The percent by mass of the solute sodium hypochlorite, NaOCl, is 3.62%. How many grams of NaOCl are in the solution?
- In question 9, how many grams of solvent are in the solution?



For more practice with percent by mass problems, go to **Supplemental Practice Problems** in Appendix A.

Percent by volume usually describes solutions in which both solute and solvent are liquids. The percent by volume is the ratio of the volume of the solute to the volume of the solution expressed as a percent. The volume of the solution is the sum of the volumes of the solute and the solvent. Calculations are similar to those involving percent by mass.

$$\text{Percent by volume} = \frac{\text{volume of solute}}{\text{volume of solution}} \times 100$$

Rubbing alcohol is an aqueous solution of liquid isopropyl alcohol. The label on a typical container, such as the one shown in **Figure 15-12**, usually states that the rubbing alcohol is 70% isopropyl alcohol. This value is a percent by volume. It tells you that 70 volumes of isopropyl alcohol are dissolved in every 100 volumes of solution. Because a solution's volume is the sum of the volumes of solute and solvent, there must be 30 volumes of water (solvent) in every 100 volumes of the rubbing alcohol.



For more practice with percent by volume problems, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

- 11.** What is the percent by volume of ethanol in a solution that contains 35 mL of ethanol dissolved in 115 mL of water?
- 12.** If you have 100.0 mL of a 30.0% aqueous solution of ethanol, what volumes of ethanol and water are in the solution?
- 13.** What is the percent by volume of isopropyl alcohol in a solution that contains 24 mL of isopropyl alcohol in 1.1 L of water?

Molarity

As you have learned, percent by volume and percent by mass are only two of the commonly used ways to quantitatively describe the concentrations of liquid solutions. One of the most common units of solution concentration is molarity. **Molarity** (M) is the number of moles of solute dissolved per liter of solution. Molarity also is known as molar concentration. The unit M is read as molar. A liter of solution containing one mole of solute is a $1M$ solution, which is read as a one molar solution. A liter of solution containing 0.1 mole of solute is a $0.1M$ solution.

To calculate a solution's molarity, you must know the volume of the solution and the amount of dissolved solute.

$$\text{Molarity } (M) = \frac{\text{moles of solute}}{\text{liters of solution}}$$

For example, suppose you need to calculate the molarity of 100.0 mL of an aqueous solution containing 0.085 mole of dissolved potassium chloride (KCl). You would first convert the volume of the solution from milliliters to liters using the conversion factor $1 \text{ L} = 1000 \text{ mL}$.

$$(100 \text{ mL}) \left(\frac{1 \text{ L}}{1000 \text{ mL}} \right) = 0.1000 \text{ L}$$

Then, to determine the molarity, you would divide the number of moles of solute by the solution volume in liters.

$$\frac{0.085 \text{ mol KCl}}{0.1000 \text{ L solution}} = \frac{0.85 \text{ mol}}{\text{L}} = 0.85M$$



Figure 15-12

The composition of this isopropyl alcohol is given in percent by volume, which is often expressed as % (v/v). What does each v in the expression 70% (v/v) refer to?



Do the **CHEMLAB** at the end of this chapter to learn about an experimental technique for determining solution concentration.

EXAMPLE PROBLEM 15-3

Calculating Molarity

A 100.5-mL intravenous (IV) solution contains 5.10 g of glucose ($\text{C}_6\text{H}_{12}\text{O}_6$). What is the molarity of this solution? The molar mass of glucose is 180.16 g/mol.

1. Analyze the Problem

You are given the mass of glucose dissolved in a volume of solution. The molarity of the solution is the ratio of moles of solute per liter of solution. Glucose is the solute and water is the solvent.

Known

mass of solute = 5.10 g $\text{C}_6\text{H}_{12}\text{O}_6$
molar mass of $\text{C}_6\text{H}_{12}\text{O}_6$ = 180.16 g/mol
volume of solution = 100.5 mL

Unknown

solution concentration = ?M

2. Solve for the Unknown

Use the molar mass to calculate the number of moles of $\text{C}_6\text{H}_{12}\text{O}_6$.

$$(5.10 \text{ g } \text{C}_6\text{H}_{12}\text{O}_6) \left(\frac{1 \text{ mol } \text{C}_6\text{H}_{12}\text{O}_6}{180.16 \text{ g } \text{C}_6\text{H}_{12}\text{O}_6} \right) = 0.0283 \text{ mol } \text{C}_6\text{H}_{12}\text{O}_6$$

Use the conversion factor $\frac{1 \text{ L}}{1000 \text{ mL}}$ to convert the volume of H_2O in milliliters to liters.

$$(100.5 \text{ mL solution}) \left(\frac{1 \text{ L}}{1000 \text{ mL}} \right) = 0.1005 \text{ L solution}$$

Substitute the known values into the equation for molarity and solve.

$$\text{molarity} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

$$\text{molarity} = \frac{0.0283 \text{ mol } \text{C}_6\text{H}_{12}\text{O}_6}{0.1005 \text{ L solution}} = \frac{0.282 \text{ mol } \text{C}_6\text{H}_{12}\text{O}_6}{\text{L solution}} = 0.282M$$

3. Evaluate the Answer

The molarity is a small value, which is expected because only a small mass of glucose was dissolved in the solution. The mass of glucose used in the problem contained three significant figures, and therefore, the value of the molarity also has three significant figures.



To prevent dehydration, intravenous (IV) drips are administered to many hospital patients. A solution containing sodium chloride and glucose is commonly used.

PRACTICE PROBLEMS

14. What is the molarity of an aqueous solution containing 40.0 g of glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) in 1.5 L of solution?
15. What is the molarity of a bleach solution containing 9.5 g of NaOCl per liter of bleach?
16. Calculate the molarity of 1.60 L of a solution containing 1.55 g of dissolved KBr.



For more practice with molarity problems, go to **Supplemental Practice Problems** in Appendix A.



Figure 15-13

Accurately preparing a solution takes care. **a** In step 1, the mass of solute to be used is measured out. **b** In step 2, the solute is placed in a volumetric flask of the correct volume. **c** In step 3, distilled water is added to the flask to bring the solution level up to the calibration mark on the flask.

Preparing Molar Solutions

Now that you know how to calculate the molarity of a solution, how do you think you would prepare 1 L of a 1.50M aqueous solution of sucrose ($C_{12}H_{22}O_{11}$) for an experiment? A 1.50M aqueous solution of sucrose contains 1.50 moles of sucrose dissolved in a liter of solution. The molar mass of sucrose is 342 g. Thus, 1.50 moles of sucrose has a mass of 513 g, an amount that you can measure on a balance.

$$\frac{1.50 \text{ mol } C_{12}H_{22}O_{11}}{1 \text{ L solution}} \times \frac{342 \text{ g } C_{12}H_{22}O_{11}}{1 \text{ mol } C_{12}H_{22}O_{11}} = \frac{513 \text{ g } C_{12}H_{22}O_{11}}{1 \text{ L solution}}$$

Unfortunately, you cannot simply add 513 g of sugar to one liter of water to make the 1.50M solution. Do you know why? Like all substances, sugar takes up space and will add volume to the solution. Therefore, you must use slightly less than one liter of water to make one liter of solution. Follow the steps shown in **Figure 15-13** to learn how to prepare the correct volume of the solution.

You often will do experiments that call for only small quantities of solution. For example, you may need only 100 mL of a 1.50M sucrose solution for an experiment. How do you determine the amount of sucrose to use? Look again at the definition of molarity. As calculated above, 1.50M solution of sucrose contains 1.50 mol of sucrose per one liter of solution. Therefore, one liter of solution contains 513 g of sucrose.

This relationship can be used as a conversion factor to calculate how much solute you need for your experiment.

$$100 \text{ mL} \times \frac{1 \cancel{\text{L}}}{1000 \cancel{\text{mL}}} \times \frac{513 \text{ g } C_{12}H_{22}O_{11}}{1 \text{ L solution}} = 51.3 \text{ g } C_{12}H_{22}O_{11}$$

Thus, you would need to measure out 51.3 g of sucrose to make 100 mL of a 1.50M solution.

PRACTICE PROBLEMS

- 17.** How many grams of $CaCl_2$ would be dissolved in 1.0 L of a 0.10M solution of $CaCl_2$?
- 18.** A liter of 2M NaOH solution contains how many grams of NaOH?
- 19.** How many grams of $CaCl_2$ should be dissolved in 500.0 mL of water to make a 0.20M solution of $CaCl_2$?
- 20.** How many grams of NaOH are in 250 mL of a 3.0M NaOH solution?

Practice!

For more practice with molarity problems, go to **Supplemental Practice Problems** in Appendix A.

Diluting solutions In the laboratory, you may use concentrated solutions of standard molarities called stock solutions. For example, concentrated hydrochloric acid (HCl) is 12*M*. Recall that a concentrated solution has a large amount of solute. You can prepare a less concentrated solution by diluting the stock solution with solvent. When you add solvent, you increase the number of solvent particles among which the solute particles move, as shown in **Figure 15-14**, thereby decreasing the solution's concentration. Would you still have the same number of moles of solute particles that were in the stock solution? Why?

How do you determine the volume of stock solution you must dilute? You know that

$$\text{Molarity (}M\text{)} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

You can rearrange the expression of molarity to solve for moles of solute.

$$\text{Moles of solute} = \text{molarity} \times \text{liters of solution}$$

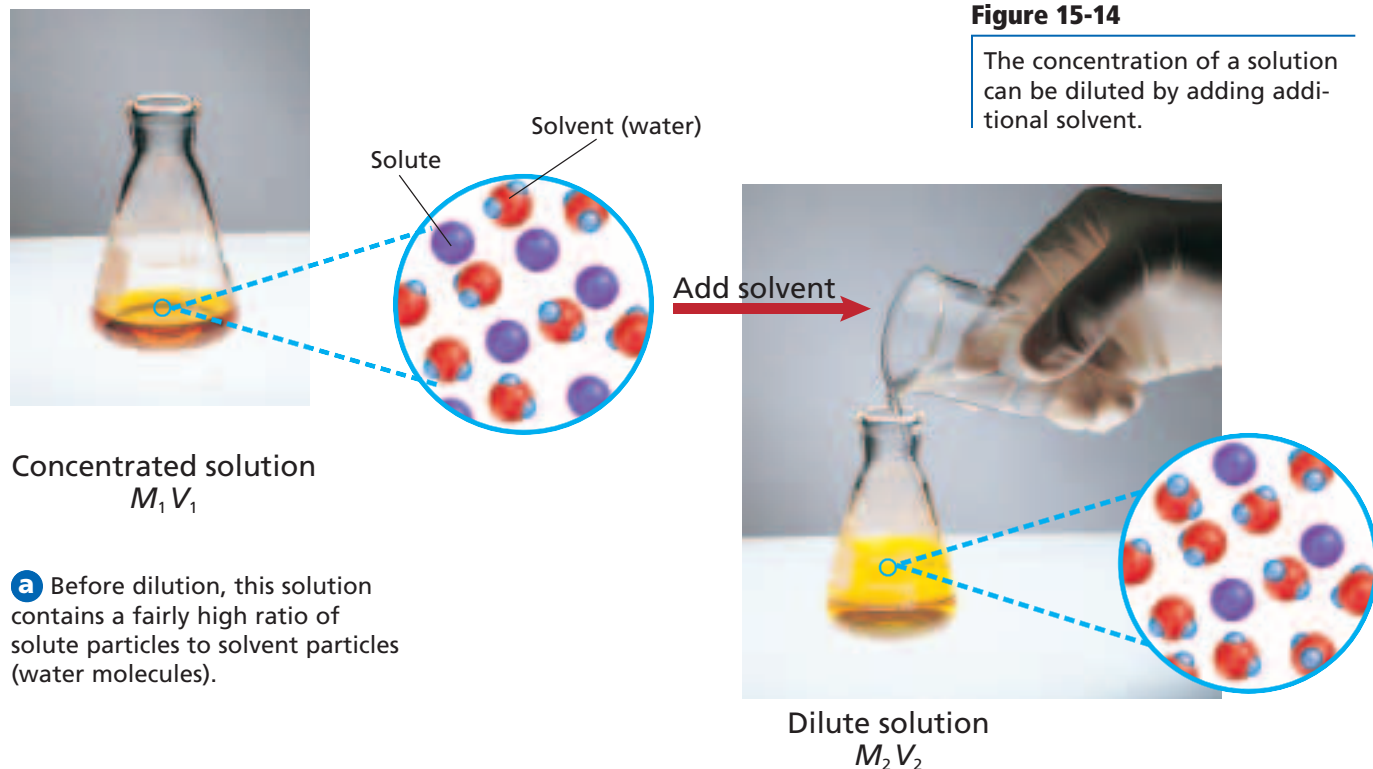
Because the total number of moles of solute does not change during dilution,

$$\text{Moles of solute in the stock solution} = \text{moles of solute after dilution}$$

You can write this relationship as the expression

$$M_1V_1 = M_2V_2$$

where M_1 and V_1 represent the molarity and volume of the stock solution and M_2 and V_2 represent the molarity and volume of the dilute solution.



EXAMPLE PROBLEM 15-4

Diluting Stock Solutions

What volume, in milliliters of 2.00M calcium chloride (CaCl_2) stock solution would you use to make 0.50 L of 0.300M calcium chloride solution?

1. Analyze the Problem

You are given the molarity of a stock solution of CaCl_2 and the volume and molarity of a dilute solution of CaCl_2 . Use the relationship between molarities and volumes to find the volume, in liters, of the stock solution required. Then convert the volume to milliliters.

Known

$$\begin{aligned}M_1 &= 2.00M \text{ CaCl}_2 \\M_2 &= 0.300M \\V_2 &= 0.50 \text{ L}\end{aligned}$$

Unknown

$$V_1 = ? \text{ L } 2.00M \text{ CaCl}_2$$



Knowing how to apply the equation $M_1V_1 = M_2V_2$ makes the preparation of dilute solutions easy.

2. Solve for the Unknown

Solve the molarity–volume relationship for the volume of the stock solution, V_1 .

$$M_1V_1 = M_2V_2$$

Dividing both sides of the equation yields,

$$V_1 = V_2 \left(\frac{M_2}{M_1} \right)$$

Substitute the known values into the equation and solve.

$$V_1 = (0.50 \text{ L}) \left(\frac{0.300M}{2.00M} \right)$$

$$V_1 = 0.075 \text{ L}$$

Use the conversion factor $\frac{1000 \text{ mL}}{1 \text{ L}}$ to convert the volume from liters to milliliters.

$$V_1 = (0.075 \cancel{\text{L}}) \left(\frac{1000 \text{ mL}}{1 \cancel{\text{L}}} \right) = 75 \text{ mL}$$

To make the dilution, measure out 75 mL of the stock solution and dilute it with enough water to make the final volume 0.50 L.

3. Evaluate the Answer

The volume V_1 was calculated and then its value was converted to milliliters. Of the given information, V_2 had the fewest number of significant figures with two. Thus, the volume V_1 should also have two significant figures, as it does.

PRACTICE PROBLEMS

21. What volume of a 3.00M KI stock solution would you use to make 0.300 L of a 1.25M KI solution?
22. How many milliliters of a 5.0M H_2SO_4 stock solution would you need to prepare 100.0 mL of 0.25M H_2SO_4 ?
23. If you dilute 20.0 mL of a 3.5M solution to make 100.0 mL of solution, what is the molarity of the dilute solution?

Practice!

For more practice with dilution problems, go to **Supplemental Practice Problems** in Appendix A.

Molality and Mole Fraction

The volume of a solution changes with temperature as it expands or contracts. This change in volume alters the molarity of the solution. Masses, however, do not change with temperature. Because of this, it is sometimes more useful to describe solutions in terms of how many moles of solute are dissolved in a specific mass of solvent. Such a description is called **molality**—the ratio of the number of moles of solute dissolved in one kilogram of solvent. The unit *m* is read as molal. A solution containing one mole of solute per kilogram of solvent is a one molal solution.

$$\text{Molality } (m) = \frac{\text{moles of solute}}{\text{kilogram of solvent}} = \frac{\text{moles of solute}}{1000 \text{ g of solvent}}$$

EXAMPLE PROBLEM 15-5

Calculating Molality

In the lab, a student adds 4.5 g of sodium chloride (NaCl) to 100.0 g of water. Calculate the molality of the solution.

1. Analyze the Problem

You are given the mass of solute and solvent. The molar mass of the solute can be used to determine the number of moles of solute in solution. Then, the molality can be calculated.

Known

mass of water (H_2O) = 100.0 g

mass of sodium chloride (NaCl) = 4.5 g

Unknown

$m = ? \text{ mol/kg}$

2. Solve for the Unknown

Use molar mass to calculate the number of moles of NaCl.

$$4.5 \text{ g NaCl} \times \frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}} = 0.077 \text{ mol NaCl}$$

Convert the mass of H_2O from grams to kilograms.

$$100.0 \text{ g H}_2\text{O} \times \frac{1 \text{ kg H}_2\text{O}}{1000 \text{ g H}_2\text{O}} = 0.1000 \text{ kg H}_2\text{O}$$

Substitute the known values into the expression for molality and solve.

$$m = \frac{\text{moles of solute}}{\text{kilogram of solvent}} = \frac{0.077 \text{ mol NaCl}}{0.1000 \text{ kg H}_2\text{O}}$$
$$m = 0.77 \text{ mol/kg}$$

3. Evaluate the Answer

Because there was less than one-tenth mole of solute present in one-tenth kilogram of water, the molality should be less than one, as it is. The mass of sodium chloride was given with two significant figures; therefore the molality also is expressed with two significant figures.



The concentration of the resulting solution can be expressed in terms of moles of solute per volume (molarity) or moles of solute per mass (molality).

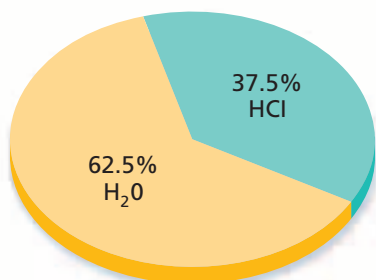
PRACTICE PROBLEMS

24. What is the molality of a solution containing 10.0 g Na_2SO_4 dissolved in 1000.0 g of water?
25. What is the molality of a solution containing 30.0 g of naphthalene (C_{10}H_8) dissolved in 500.0 g of toluene?



For more practice with molality problems, go to **Supplemental Practice Problems** in Appendix A.

Hydrochloric Acid in Aqueous Solution



$$X_{\text{HCl}} + X_{\text{H}_2\text{O}} = 1.000$$

$$0.229 + 0.771 = 1.000$$

Figure 15-15

The mole fraction expresses the number of moles of solvent and solute relative to the total number of moles of solution. Each mole fraction can be thought of as a percent. For example, the mole fraction of water ($X_{\text{H}_2\text{O}}$) is 0.771, which is equivalent to saying the solution contains 77.1% water (on a mole basis).

Mole fraction If you know the number of moles of solute and solvent, you can also express the concentration of a solution in what is known as a **mole fraction**—the ratio of the number of moles of solute in solution to the total number of moles of solute and solvent.

The symbol X is commonly used for mole fraction, with a subscript to indicate the solvent or solute. The mole fraction for the solvent (X_A) and the mole fraction for the solute (X_B) can be expressed as follows

$$X_A = \frac{n_A}{n_A + n_B} \quad X_B = \frac{n_B}{n_A + n_B}$$

where n_A is the number of moles of solvent and n_B is the number of moles of solute. Why must the sum of the mole fractions for all components in a solution equal one?

Consider as an example the mole fraction of hydrochloric acid (HCl) in the aqueous solution shown in **Figure 15-15**. For every 100 grams of solution, 37.5 g would be HCl and 62.5 g would be H_2O . To convert these masses to moles, you would use the molar masses as conversion factors.

$$n_{\text{HCl}} = 37.5 \text{ g HCl} \times \frac{1 \text{ mol HCl}}{36.5 \text{ g HCl}} = 1.03 \text{ mol HCl}$$

$$n_{\text{H}_2\text{O}} = 62.5 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.0 \text{ g H}_2\text{O}} = 3.47 \text{ mol H}_2\text{O}$$

Thus, the mole fractions of hydrochloric acid and water can be expressed as

$$X_{\text{HCl}} = \frac{n_{\text{HCl}}}{n_{\text{HCl}} + n_{\text{H}_2\text{O}}} = \frac{1.03 \text{ mol HCl}}{1.03 \text{ mol HCl} + 3.47 \text{ mol H}_2\text{O}}$$

$$X_{\text{HCl}} = 0.229$$

$$X_{\text{H}_2\text{O}} = \frac{n_{\text{H}_2\text{O}}}{n_{\text{HCl}} + n_{\text{H}_2\text{O}}} = \frac{3.47 \text{ mol H}_2\text{O}}{1.03 \text{ mol HCl} + 3.47 \text{ mol H}_2\text{O}}$$

$$X_{\text{H}_2\text{O}} = 0.771$$

PRACTICE PROBLEMS

- 26.** What is the mole fraction of NaOH in an aqueous solution that contains 22.8% NaOH by mass?
- 27.** An aqueous solution of NaCl has a mole fraction of 0.21. What is the mass of NaCl dissolved in 100.0 mL of solution?



For more practice with mole fraction problems, go to **Supplemental Problems** in Appendix A.

Section 15.2 Assessment

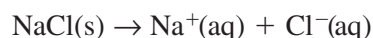
- 28.** Distinguish between a dilute solution and a concentrated solution.
- 29.** Compare and contrast five quantitative ways to describe the composition of solutions.
- 30.** Describe the laboratory procedure for preparing a specific volume of a dilute solution from a concentrated stock solution.
- 31. Thinking Critically** Explain the similarities and differences between a 1M solution of NaOH and a 1m solution of NaOH.
- 32. Using Numbers** A can of chicken broth contains 450 mg of sodium chloride in 240.0 g of broth. What is the percent by mass of sodium chloride in the broth?

Solutes affect some of the physical properties of their solvents. Early researchers were puzzled to discover that the effects of a solute on a solvent depended only on how many solute particles were in solution, not on the specific solute dissolved. Physical properties of solutions that are affected by the number of particles but not the identity of dissolved solute particles are called **colligative properties**. The word colligative means “depending on the collection.” Colligative properties include vapor pressure lowering, boiling point elevation, freezing point depression, and osmotic pressure.

Electrolytes and Colligative Properties

In Chapter 8, you learned that ionic compounds are called electrolytes because they dissociate in water to form a solution that conducts electric current. Some molecular compounds ionize in water and also are electrolytes. Electrolytes that produce many ions in solution are called strong electrolytes; those that produce only a few ions in solution are called weak electrolytes.

Sodium chloride is a strong electrolyte. It almost completely dissociates in solution, producing Na^+ and Cl^- ions.



Dissolving one mole of NaCl in a kilogram of water does not yield a 1*m* solution of ions. Rather, there would be almost two moles of solute particles in solution—approximately one mole each of Na^+ and Cl^- ions. How many moles of ions would you expect to find in a 1*m* aqueous solution of HCl?

Nonelectrolytes in aqueous solution Many molecular compounds dissolve in solvents but do not ionize. Such solutions do not conduct an electric current, and the solutes are called nonelectrolytes. Sucrose is an example of a nonelectrolyte. A 1*m* sucrose solution contains only one mole of sucrose particles. **Figure 15-16** compares the conductivity of a solution containing an electrolyte solute with one containing a nonelectrolyte solute. Which compound would have the greater effect on colligative properties, sodium chloride or sucrose?

Objectives

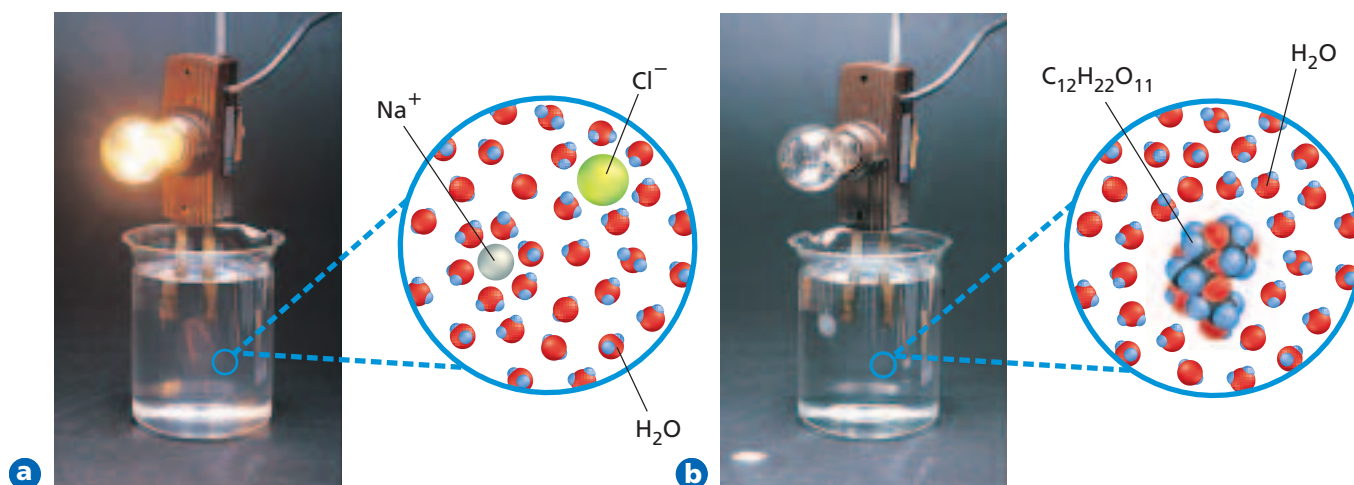
- **Explain** the nature of colligative properties.
- **Describe** four colligative properties of solutions.
- **Calculate** the boiling point elevation and the freezing point depression of a solution.

Vocabulary

colligative property
vapor pressure lowering
boiling point elevation
freezing point depression
osmosis
osmotic pressure

Figure 15-16

a Sodium chloride is a strong electrolyte and conducts electricity well. **b** Sucrose, while soluble in water, does not ionize and therefore does not conduct electricity. Which solute, sodium chloride or sucrose, produces more particles in solution per mole?



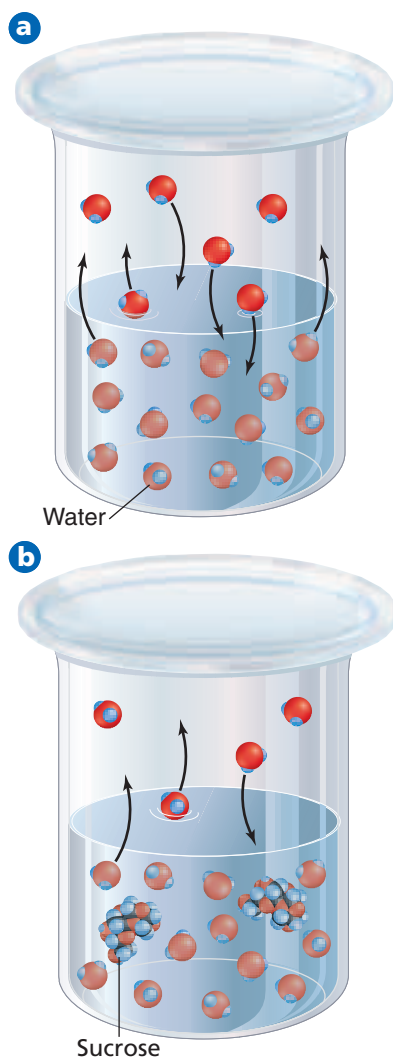


Figure 15-17

The vapor pressure of a pure solvent **a** is greater than the vapor pressure of a solution containing a nonvolatile solute **b**.

Vapor Pressure Lowering

In Chapter 13, you learned that vapor pressure is the pressure exerted in a closed container by liquid particles that have escaped the liquid's surface and entered the gaseous state. In a closed container at constant temperature and pressure, the solvent particles reach a state of dynamic equilibrium, escaping and reentering the liquid state at the same rate. The vapor pressure for a closed container of pure water is shown in **Figure 15-17a**.

Experiments show that adding a nonvolatile solute (one that has little tendency to become a gas) to a solvent lowers the solvent's vapor pressure. The particles that produce vapor pressure escape the liquid phase at its surface. When a solvent is pure, as shown in **Figure 15-17a**, its particles occupy the entire surface area. However, when the solvent contains solute, as shown in **Figure 15-17b**, a mix of solute and solvent particles occupies the surface area. With fewer solvent particles at the surface, fewer particles enter the gaseous state, and the vapor pressure is lowered. The greater the number of solute particles in a solvent, the lower the resulting vapor pressure. Thus, **vapor pressure lowering** is due to the number of solute particles in solution and is a colligative property of solutions.

You can predict the relative effect of a solute on vapor pressure based on whether the solute is an electrolyte or a nonelectrolyte. For example, one mole each of the solvated nonelectrolytes glucose, sucrose, and ethanol molecules has the same relative effect on the vapor pressure. However, one mole each of the solvated electrolytes sodium chloride, sodium sulfate, and aluminum chloride has an increasingly greater effect on vapor pressure because of the increasing number of ions each produces in solution.

Boiling Point Elevation

Because a solute lowers a solvent's vapor pressure, it also affects the boiling point of the solvent. Recall from Chapter 13 that the liquid in a pot on your stove boils when its vapor pressure equals atmospheric pressure. When the temperature of a solution containing a nonvolatile solute is raised to the boiling point of the pure solvent, the resulting vapor pressure is still less than the atmospheric pressure and the solution will not boil. Thus, the solution must be heated to a higher temperature to supply the additional kinetic energy needed to raise the vapor pressure to atmospheric pressure. The temperature difference between a solution's boiling point and a pure solvent's boiling point is called the **boiling point elevation**.

For nonelectrolytes, the value of the boiling point elevation, which is symbolized ΔT_b , is directly proportional to the solution's molality.

$$\Delta T_b = K_b m$$

The molal boiling point elevation constant, K_b , is the difference in boiling points between a $1m$ nonvolatile, nonelectrolyte solution and a pure solvent. It is expressed in units of $^{\circ}\text{C}/m$ and varies for different solvents. Values of K_b for several common solvents are found in **Table 15-4**. Note that water's K_b value is $0.512^{\circ}\text{C}/m$. This means that a $1m$ aqueous solution containing a nonvolatile, nonelectrolyte solute boils at 100.512°C , a temperature 0.512°C higher than pure water's boiling point of 100.0°C .

Like vapor pressure lowering, boiling point elevation is a colligative property. The value of the boiling point elevation is directly proportional to the solution's solute molality, that is, the greater the number of solute particles in the solution, the greater the boiling point elevation.

Table 15-4

Molal Boiling-Point Elevation Constants (K_b)		
Solvent	Boiling point ($^{\circ}\text{C}$)	K_b ($^{\circ}\text{C}/m$)
Water	100.0	0.512
Benzene	80.1	2.53
Carbon tetrachloride	76.7	5.03
Ethanol	78.5	1.22
Chloroform	61.7	3.63

Freezing Point Depression

The freezing point depression of a solution is another colligative property of solutions. At a solvent's freezing point temperature, the particles no longer have sufficient kinetic energy to overcome the interparticle attractive forces; the particles form into a more organized structure in the solid state. In a solution, the solute particles interfere with the attractive forces among the solvent particles. This prevents the solvent from entering the solid state at its normal freezing point. The freezing point of a solution is always lower than that of a pure solvent. **Figure 15-18** shows the differences in boiling and melting points of pure water and an aqueous solution. By comparing the solid and dashed lines, you can see that the temperature range over which the aqueous solution exists as a liquid is greater than that of pure water. You can observe the effect of freezing point depression by doing the **miniLAB** on this page.

A solution's **freezing point depression**, ΔT_f , is the difference in temperature between its freezing point and the freezing point of its pure solvent. For nonelectrolytes, the value of the freezing point depression, which is symbolized as ΔT_f , is directly proportional to the solution's molality.

$$\Delta T_f = K_f m$$

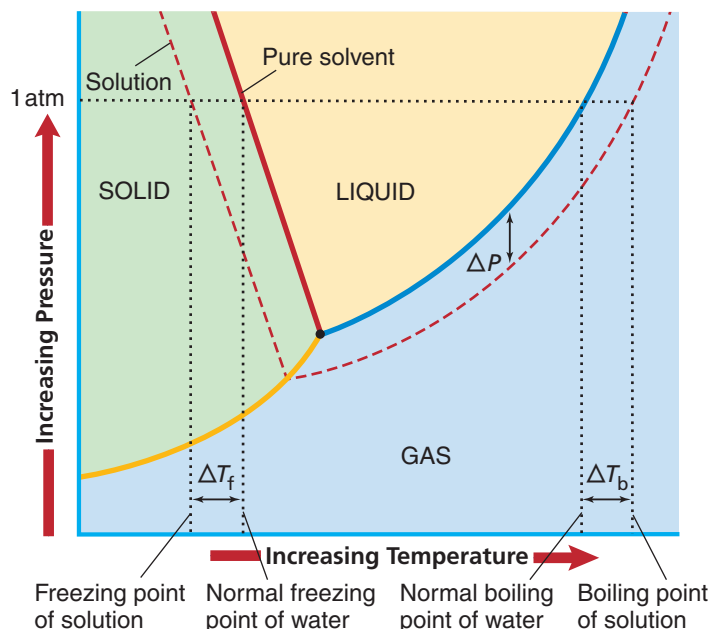


Figure 15-18

This phase diagram shows how temperature and pressure affect the solid, liquid, and gas phases of a pure solvent (solid lines) and a solution (dashed lines). The difference between the solid and dashed lines corresponds to vapor pressure lowering (ΔP), boiling point elevation (ΔT_b), and freezing point depression (ΔT_f).

miniLAB

Freezing Point Depression

Measuring The colligative property of freezing point depression can be observed in a simple laboratory investigation. You will measure the temperatures of two beakers and their contents.

Materials 400-mL beakers (2), crushed ice, rock salt (NaCl), water, stirring rods (2), graduated cylinder, thermometers (2), balance

Procedure

1. Fill two 400-mL beakers with crushed ice. Add 50 mL of cold tap water to each beaker.
2. Stir the contents of each beaker with a stirring rod until both beakers are at a constant temperature, approximately one minute.
3. Measure the temperature of each beaker using a thermometer and record the readings.
4. Add 75 g of rock salt to one of the beakers. Continue stirring both beakers. Some of the salt will dissolve.

5. When the temperature in each beaker is constant, record the readings.
6. To clean up, flush the contents of each beaker down the drain with excess water.

Analysis

1. Compare your readings taken for the ice water and the salt water. How do you explain the observed temperature change?
2. Why was salt only added to one of the beakers?
3. Salt is a strong electrolyte that produces two ions, Na^+ and Cl^- , when it dissociates in water. Why is this important to consider when calculating the colligative property of freezing point depression?
4. Predict if it would be better to use coarse rock salt or fine table salt when making homemade ice cream. Explain.

Values of K_f for several common solvents are found in **Table 15-5**. As with K_b values, K_f values are specific to their solvents. With water's K_f value of $1.86^\circ\text{C}/m$, a $1m$ aqueous solution containing a nonvolatile, nonelectrolyte solute freezes at -1.86°C rather than at pure water's freezing point of 0.0°C .

However, in using the ΔT_b and ΔT_f relationships with electrolytes, you must make sure to use the effective molality of the solution. For example, one mole of the electrolyte sodium chloride (NaCl) forms one mole of Na^+ ions and one mole of Cl^- ions in solution. Thus, a $1m$ aqueous NaCl solution produces two moles of solute particles in a kilogram of water—effectively acting as a $2m$ solution. Because of this, $2m$, and not $1m$, must be used in the ΔT_b and ΔT_f relationships for a $1m$ solution of sodium chloride. The following Example Problem illustrates this point.

Table 15-5

Molal Freezing Point Depression Constants (K_f)		
Solvent	Freezing point ($^\circ\text{C}$)	K_f ($^\circ\text{C}/m$)
Water	0.0	1.86
Benzene	5.5	5.12
Carbon tetrachloride	-23	29.8
Ethanol	-114.1	1.99
Chloroform	-63.5	4.68

EXAMPLE PROBLEM 15-6

Changes in Boiling and Freezing Points

What are the boiling point and freezing point of a $0.029m$ aqueous solution of sodium chloride (NaCl)?

1. Analyze the Problem

You are given the molality of an aqueous sodium chloride solution. First, calculate ΔT_b and ΔT_f based on the number of particles in solution. Then, to determine the elevated boiling point and the depressed freezing point, add ΔT_b to the normal boiling point and subtract ΔT_f from the normal freezing point.

Known

solute = sodium chloride (NaCl)

molality of sodium chloride solution = $0.029m$

Unknown

boiling point = ? $^\circ\text{C}$

freezing point = ? $^\circ\text{C}$

2. Solve for the Unknown

Each mole of the electrolyte sodium chloride dissociates in solution to produce two moles of particles. Calculate the effective number of solute particles in solution.

$$\text{particle molality} = 2 \times 0.029m = 0.058m$$

Substitute the known values for K_b , K_f , and particle molality into the ΔT_b and ΔT_f equations and solve.

$$\Delta T_b = K_b m = (0.512^\circ\text{C}/m)(0.058m) = 0.030^\circ\text{C}$$

$$\Delta T_f = K_f m = (1.86^\circ\text{C}/m)(0.058m) = 0.11^\circ\text{C}$$

Add ΔT_b to the normal boiling point and subtract ΔT_f from the normal freezing point to determine the elevated boiling point and depressed freezing point of the solution.

$$\text{boiling point} = 100.0^\circ\text{C} + 0.030^\circ\text{C} = 100.030^\circ\text{C}$$

$$\text{freezing point} = 0.0^\circ\text{C} - 0.11^\circ\text{C} = -0.11^\circ\text{C}$$

3. Evaluate the Answer

The boiling point is higher and the freezing point is lower, as expected. Because the molality of the solution has two significant figures, both ΔT_b and ΔT_f also have two significant figures. Because the normal boiling point and freezing point are exact values, they do not affect the number of significant figures in the final answer.

PRACTICE PROBLEMS

33. What are the boiling point and freezing point of a 0.625*m* aqueous solution of any nonvolatile, nonelectrolyte solute?
34. What are the boiling point and freezing point of a 0.40*m* solution of sucrose in ethanol?
35. A lab technician determines the boiling point elevation of an aqueous solution of a nonvolatile, nonelectrolyte to be 1.12°C. What is the solution's molality?
36. A student dissolves 0.500 mol of a nonvolatile, nonelectrolyte solute in one kilogram of benzene (C₆H₆). What is the boiling point elevation of the resulting solution?



For more practice with boiling point elevation and freezing point depression problems, go to **Supplemental Practice Problems** in Appendix A.

Osmosis and Osmotic Pressure

In Chapter 13, you learned that diffusion is the mixing of gases or liquids resulting from their random motions. **Osmosis** is the diffusion of solvent particles across a semipermeable membrane from an area of higher solvent concentration to an area of lower solvent concentration. Semipermeable membranes are barriers with tiny pores that allow some but not all kinds of particles to cross. The membranes surrounding all living cells are semipermeable membranes. Osmosis plays an important role in many biological systems such as kidney dialysis and the uptake of nutrients by plants.

Let's look at a simple system in which a sucrose-water solution is separated from its solvent—pure water—by a semipermeable membrane. During osmosis, water molecules move in both directions across the membrane, but the sugar molecules cannot cross it. Both water and sugar molecules contact the membrane on the solution side, but only water molecules contact the membrane on the pure solvent side. Thus, more water molecules cross the membrane from the pure solvent side than from the solution side.

The additional water molecules on the solution side of the membrane create pressure and push some water molecules back across the membrane. The amount of additional pressure caused by the water molecules that moved into the solution is called the **osmotic pressure**. Osmotic pressure depends upon the number of solute particles in a given volume of solution. Therefore, osmotic pressure is another colligative property of solutions.

Careers Using Chemistry

Renal Dialysis Technician

People with poorly functioning kidneys must routinely undergo dialysis in order to survive. The process, which removes impurities from the blood through osmosis, can be stressful for the patient. If you enjoy working with and helping people, and want to work in a medical setting, consider being a renal dialysis technician.

Dialysis technicians prepare the dialysis equipment and the patient for treatment, start the treatment, and monitor the process. They also make necessary adjustments, keep records, and respond to emergencies. Technicians spend more time with the patients than the doctors or nurses. They work in hospitals, outpatient facilities, and home-based dialysis programs.

Section 15.3 Assessment

37. Explain the nature of colligative properties.
38. Describe four colligative properties of solutions.
39. Explain why a solution has a lower boiling point than the pure solvent.
40. **Thinking Critically** Explain why the colligative properties described in this section may not apply to solutions containing volatile solutes. Hint: Volatile solutes are able to leave the liquid phase and enter the gas phase.
41. **Using Numbers** Calculate the boiling point elevation and freezing point depression of a solution containing 50.0 g of glucose (C₆H₁₂O₆) dissolved in 500.0 g of water.



Objectives

- **Identify** the properties of suspensions and colloids.
- **Describe** different types of colloids.
- **Explain** the electrostatic forces in colloids.

Vocabulary

suspension
colloid
Brownian motion
Tyndall effect

As you learned in Chapter 3, most of the forms of matter that you encounter are mixtures. A mixture is a combination of two or more substances that keep their basic identity. Components of a mixture come in contact with each other but do not undergo chemical change. You have been studying homogeneous mixtures called solutions so far in this chapter. Not all mixtures are solutions, however. Heterogeneous mixtures contain substances that exist in distinct phases. Two types of heterogeneous mixtures are suspensions and colloids.

Suspensions

Look at the mixture shown in **Figure 15-19a**. Although it resembles milk, it is actually a freshly-made mixture of cornstarch stirred into water. If you let this mixture stand undisturbed for a while, it separates into two distinct layers: a thick, white, pastelike substance on the bottom and the water on top. Cornstarch in water is a **suspension**, a mixture containing particles that settle out if left undisturbed. As shown in **Figure 15-19b**, pouring a liquid suspension through a filter also separates out the suspended particles. Other examples of suspensions include fine sand in water and muddy water.

Suspended particles are large compared to solvated particles, with diameters greater than 1000 nm (10^{-6} m) compared to diameters less than 1 nm (10^{-9} m) for solvated particles. Gravity acts on suspended particles in a short time, causing them to settle out of the mixture. Interestingly, the settled out cornstarch particles form a solidlike state on the bottom of the container. However, when stirred, the solidlike state quickly begins flowing like a liquid. Substances that behave this way are called thixotropic. One of the most common applications of a thixotropic mixture is house paint. The paint flows rather easily when applied with a brush, but quickly thickens to a solidlike state. The quickly thickening paint helps it stick to the house and prevents runs in the paint.

Figure 15-19

A suspension is a type of heterogeneous mixture. Suspension particles settle out over time **a**, and can be separated from the mixture by filtration **b**.




Table 15-6

Types of Colloids			
Category	Dispersed particles	Dispersing medium	Example
Solid sol	Solid	Solid	Colored gems
Sol	Solid	Liquid	Blood, gelatin
Solid emulsion	Liquid	Solid	Butter, cheese
Emulsion	Liquid	Liquid	Milk, mayonnaise
Solid foam	Gas	Solid	Marshmallow, soaps that float
Foam	Gas	Liquid	Whipped cream, beaten egg white
Aerosol	Solid	Gas	Smoke, dust in air
Aerosol	Liquid	Gas	Spray deodorant, clouds

Colloids

Particles in a suspension are much larger than atoms. In contrast, particles in a solution are atomic-scale in size. A heterogeneous mixture of intermediate size particles (between the size of solution particles and suspension particles) is a **colloid**. Colloid particles are between 1 nm and 1000 nm in diameter. The most abundant substance in the mixture is the dispersion medium. Milk is a colloid. Although homogenized milk resembles the cornstarch mixture in **Figure 15-19a**, you cannot separate its components by settling or by filtration.

Colloids are categorized according to the phases of their dispersed particles and dispersing mediums. Milk is a colloidal emulsion because liquid particles are dispersed in a liquid medium. Other types of colloids are described in **Table 15-6** and shown in **Figure 15-20**. How many of them are familiar to you? Can you name others?


LAB

See page 959 in Appendix E for
Identifying Colloids

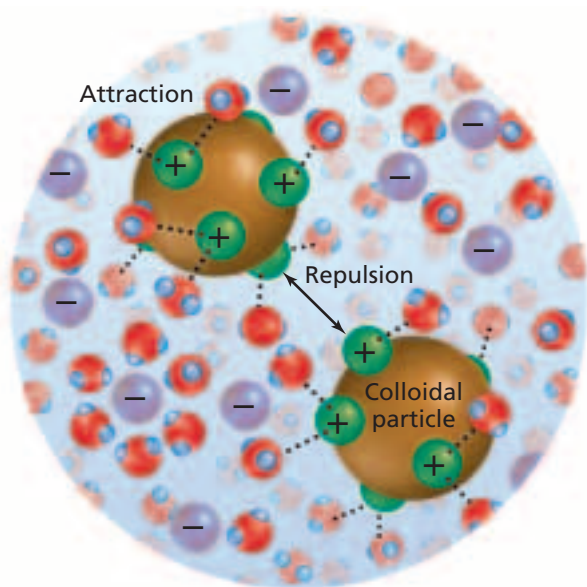
Figure 15-20

All of the photos shown here are examples of colloids.
a Fog, which is a liquid aerosol, is formed when small liquid particles are dispersed in a gas. **b** Many paints are sols, a fluid colloidal system of fine solid particles in a liquid medium. **c** The bagel has a coarse texture with lots of small holes throughout. These small holes are formed by a type of foam produced by yeast in the dough.



Figure 15-21

The dispersing medium particles form charged layers around the colloid particles. These charged layers repel each other and keep the particles from settling out.



Brownian motion If you were to observe a liquid colloid under the magnification of a microscope, you would see that the dispersed particles make jerky, random movements. This erratic movement of colloid particles is called **Brownian motion**. It was first observed by and later named for the Scottish botanist Robert Brown, who noticed the random movements of pollen grains dispersed in water. Brownian motion results from collisions of particles of the dispersion medium with the dispersed particles. These collisions prevent the colloid particles from settling out of the mixture.

The reason the dispersed particles do not settle out is because they have polar or charged atomic groups on their surfaces. These areas on their surfaces attract the positively or negatively charged areas of the dispersing-medium particles. This results in the formation of electrostatic layers around the particles, as shown in **Figure 15-21**. The layers repel each other when the dispersed particles collide and, thus, the particles remain in the colloid. If you interfere with the electrostatic layering, colloid particles will settle out of the mixture. For example, if you stir an electrolyte into a colloid, the dispersed particles increase in size, destroying the colloid. Heating also destroys a colloid because it gives colliding particles enough kinetic energy to overcome the electrostatic forces and settle out.

problem-solving LAB

How can you measure the turbidity of a colloid?

Designing an Experiment The presence of cloudy water, or *turbidity*, in a lake or river results from the suspension of solids in the water. Most of these colloid particles come from erosion, industrial and human waste, algae blooms from fertilizers, and decaying organic matter. The degree of turbidity can greatly affect the health of a body of water in a number of ways, but its greatest impact is that it prevents light from reaching vegetation.

Analysis

The Tyndall effect can be used to measure the turbidity of water. Your goal is to plan a procedure and develop a scale to interpret the data.

Thinking Critically

Your plan should address the following.

1. What materials would you need to measure the Tyndall effect? Can you select CBLs, computers, or graphing calculators to collect or interpret data?
2. What variables can be used to relate the ability of light to pass through the liquid and the number of the colloid particles present?
3. How would you change the turbidity in the experiment?
4. What variable would change and by how much?
5. What will you use as a control?
6. How could you relate the variables used in the experiment to the actual number of colloid particles that are present?
7. Are there any safety precautions you must consider?



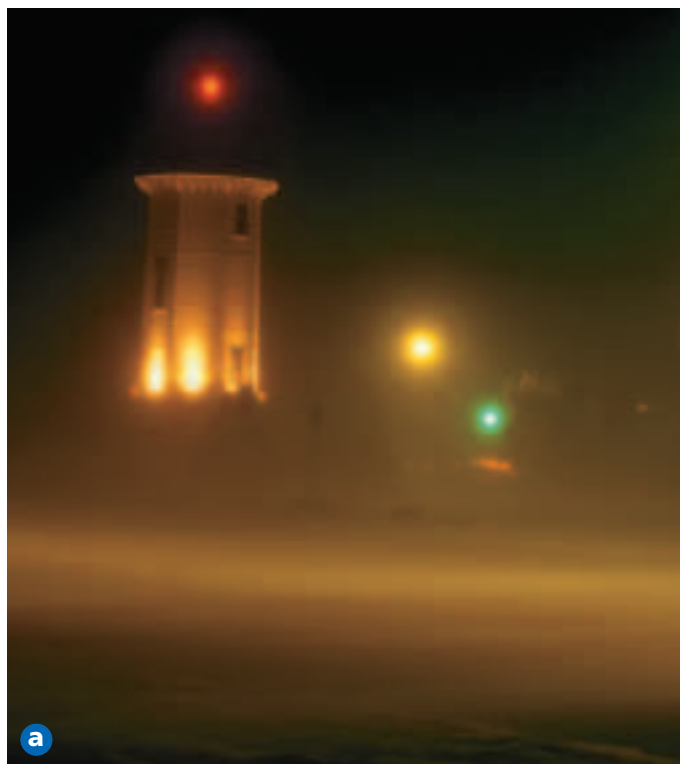


Figure 15-22

Two examples of the Tyndall effect are shown here. **a** The scattering of light by fog greatly reduces visibility. **b** The beam of light is visible in the colloid because of light scattering.

The Tyndall effect Whereas concentrated colloids are often cloudy or opaque, dilute colloids sometimes appear as clear as solutions. Dilute colloids appear to be homogeneous because their dispersed particles are so small that they cannot be seen by the unaided eye. However, dispersed colloid particles are large enough to scatter light, a phenomenon known as the **Tyndall effect**. In **Figure 15-22b**, a beam of light is shone through two unknown mixtures. Which mixture is the colloid? Which is the solution? You can see that dispersed colloid particles scatter the light, unlike solvated particles in the solution. Solutions never exhibit the Tyndall effect. Suspensions, such as the cornstarch shown in **Figure 15-19a**, also exhibit the Tyndall effect. You have observed the Tyndall effect if you have observed rays of sunlight passing through smoke-filled air, or viewed lights through fog at night, as shown in **Figure 15-22a**. Do the **problem-solving LAB** on the previous page to see how the Tyndall effect can be used to determine the amount of colloid particles in suspension.

Section 15.4 Assessment

- 42.** Distinguish between suspensions and colloids.
- 43.** Describe different types of colloids.
- 44.** Why do dispersed colloid particles stay dispersed?
- 45. Thinking Critically** Use the Tyndall effect to explain why it is more difficult to drive through fog using high beams than using low beams.
- 46. Comparing and Contrasting** Make a table that compares the properties of solutions, suspensions, and colloids.



Beer's Law

Finding the concentration of an unknown solution is an important procedure in laboratory work. One method commonly used to determine solution concentration is to measure how much of a single wavelength of light is absorbed by the solution and compare it to known values of concentration and wavelength. Light absorbance is directly related to the concentration of a solution. This relationship is called Beer's law.

Problem

How is light absorbance used to find the concentration of a blue dye solution?

Objectives

- **Prepare** solutions of known concentration from a blue dye stock solution.
- **Measure** the absorbance of known and unknown aqueous solutions.
- **Infer** the relationship between light absorbance and concentration of a solution.

Materials

CBL system
graphing calculator
Vernier colorimeter
DIN adaptor and cable
TI Graph-Link (optional)
cuvette
cotton swabs
tissues for wiping
cuvette

blue dye stock solution
unknown solution
distilled water
50-mL graduated cylinder
100-mL beaker (2)
small test tube (5)
test tube rack
pipette (2)
pipette bulb
stirring rod

Safety Precautions



- Always wear safety goggles and a lab apron.
- The food-coloring solution can stain clothes.

Pre-Lab

1. Read the entire **CHEMLAB** procedure.
2. What is the total volume of solution in each test tube? Calculate the percent by volume of the solutions in test tubes 1 through 5. Prepare a data table.
3. Review with your teacher how a colorimeter works. How are absorbance (A) and transmittance ($%T$) related?

Known and Unknown Solutions Data			
Test tube	Concentration (%)	$%T$	A
1			
2			
3			
4			
5			
Unknown			

4. What is occurring during step 3 of the procedure? Why is a cuvette of water used?

Procedure

1. Transfer 30 mL of blue dye stock solution into a beaker. Transfer 30 mL of distilled water into another beaker.
2. Label five clean, dry test tubes 1 through 5.
3. Pipette 2 mL of blue dye stock solution from the beaker into test tube 1, 4 mL into test tube 2, 6 mL into test tube 3, and 8 mL into test tube 4.
CAUTION: Always pipette using a pipette bulb.
4. With another pipette, transfer 8 mL of distilled water from the beaker into test tube 1, 6 mL into tube 2, 4 mL into tube 3, and 2 mL into tube 4.
5. Mix the solution in test tube 1 with a stirring rod. Rinse and dry the stirring rod. Repeat this procedure with each test tube.

6. Pipette 10 mL of blue dye stock into test tube 5.
7. Load the ChemBio program into the calculator. Connect the CBL to the colorimeter using a DIN adaptor. Connect the CBL to the calculator using a link cable. Begin the ChemBio program on the calculator. Select "1" probe. Select 4: COLORIMETER. Enter Channel "1."
8. Fill a cuvette about three-fourths full with distilled water and dry its outside with a tissue. To calibrate the colorimeter, place the cuvette in the colorimeter and close the lid. Turn the wavelength knob to 0%T. Press TRIGGER on the CBL and enter 0 into the calculator. Turn the wavelength knob to Red (635 nm). Press TRIGGER on the CBL and enter 100 into the calculator. Leave the colorimeter set on Red for the rest of the lab. Remove the cuvette from the colorimeter. Empty the distilled water from the cuvette. Dry the inside of the cuvette with a clean cotton swab.
9. Select COLLECT DATA from the MAIN MENU. Select TRIGGER/PROMPT from the DATA COLLECTION menu. Fill the cuvette about three-fourths full with the solution from test tube 1. Dry the outside of the cuvette with a tissue and place the cuvette in the colorimeter. Close the lid. After 10 to 15 seconds, press TRIGGER and enter the concentration in percent from your data table into the calculator. Remove the cuvette and pour out the solution. Rinse the inside of the cuvette with distilled water and dry it with a clean cotton swab. Repeat this step for test tubes 2 through 5.
10. Select STOP AND GRAPH from the DATA COLLECTION menu when you have finished with data collection. Draw the graph, or use the TI Graph-Link to make a copy of the graph from the calculator screen. You also will want to copy the data from the STAT list into your data table (or you can print it from a screen print using Graph-Link).
11. Clean the cuvette with a cotton swab and fill it about three-fourths full with the unknown dye solution. Place the cuvette in the colorimeter and close the lid. From the MAIN MENU, select COLLECT DATA (do not select SET UP PROBES as this will erase your data lists). Select

MONITOR INPUT from the DATA COLLECTION MENU. Press ENTER to monitor the absorbance value of the colorimeter. After about 10-15 seconds, record the absorbance value and record it in your data table.

Cleanup and Disposal

1. All of the blue dye solutions can be rinsed down the drain.
2. Turn off the colorimeter. Clean and dry the cuvette. Return all equipment to its proper place.

Analyze and Conclude

1. **Analyzing Data** Evaluate how close your graph is to the direct relationship exhibited by Beer's law by doing a linear-regression line. Select FIT CURVE from the MAIN MENU (do not select SET UP PROBES as this will erase your data lists). Select LINEAR L1,L2. The calculator will give you an equation in the form of $y = ax + b$. One indicator of the fit of your graph is the size of b . A small value of b means the graph passes close to the origin. The closer the correlation coefficient r reported by the program is to 1.00, the better the fit of the graph.
2. **Drawing Conclusions** Use the graph of your absorbance and concentration data to determine the concentration of your unknown solution.
3. **Form a Hypothesis** Would you obtain the same data if red dye was used? Explain.
4. **Error Analysis** Analyze your b and r values. How closely do your results match Beer's law? Reexamine the procedure and suggest reasons why the correlation coefficient from your data does not equal 1.00.

Real-World Chemistry

1. Explain how Beer's law can be applied in food, drug, and medical testing.
2. The reaction of alcohol with orange dichromate ions to produce blue-green chromium(III) ions is used in the Breathalyzer test, a test that measures the presence of alcohol in a person's breath. How could a colorimeter be used in this analysis?

CHEMISTRY and Society

Sickle Cell Disease

There are millions of red blood cells in a single drop of blood. Red blood cells play a crucial role, transporting oxygen throughout the body.

Hemoglobin in Red Blood Cells

The cells pictured at the right appear red because they contain a bright red molecule—a protein called hemoglobin. One molecule of hemoglobin contains four iron(II) ions. Four oxygen molecules bind to each molecule of hemoglobin as blood passes through the lungs. As the blood circulates through the body, hemoglobin releases oxygen to the cells that make up various body tissues.

Healthy red blood cells are round and flexible, passing easily through narrow capillaries. But, hemoglobin in people with sickle cell disease differs from normal hemoglobin.

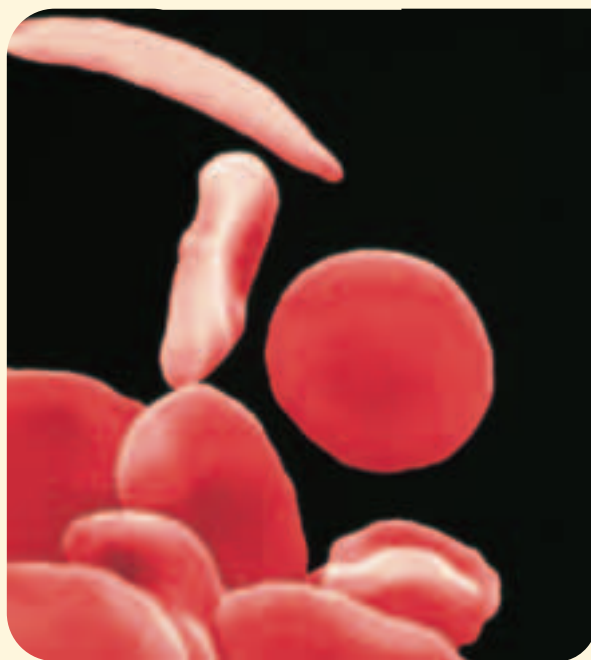
What makes sickle cells different?

Each hemoglobin molecule consists of four chains. The entire molecule contains 574 amino acid groups. In people with sickle cell disease, two non-polar amino acid groups are substituted for two polar amino acid groups. As a result, there is less repulsion between hemoglobin molecules, which allows abnormal hemoglobin molecules to clump together. This causes the abnormal hemoglobin molecules to be less soluble in the red blood cells, especially after oxygen has been absorbed through the lungs. The abnormal hemoglobin forces the red blood cells to become rigid and C-shaped, resembling the farming tool called a sickle.

These unusually shaped sickle cells clog the circulatory system, reducing blood flow. Therefore, oxygen supply to nearby tissues is reduced. Sickle cell disease causes pain, anemia, stroke, and susceptibility to infection.

Searching for a Cure

While most individuals with sickle cell disease have Mediterranean or African ancestry, it is common practice to screen all newborns in the United States for the condition. Research shows that early intervention can reduce the risk of serious infection, the leading cause of death in children with sickle cell disease. Intensive chemotherapy and stem cell



transfusions are being studied as ways to treat and possibly cure sickle cell disease.

Investigating the Issue

- 1. Communicating Ideas** Learn about the latest advances in sickle cell disease research. Create a chart that shows major symptoms, their causes, and their treatment. Share your findings with other students in your class.
- 2. Debating the Issue** The intensive chemotherapy required to treat sickle cell disease drastically weakens the immune system, leaving the patient vulnerable to overwhelming infection. Five to eight percent of the children who undergo this treatment do not survive. Should doctors be allowed to use this procedure to treat sickle cell disease?



Visit the Chemistry Web site at chemistrymc.com to find links to more information about sickle cell disease.

Summary

15.1 What are solutions?

- A solute dissolves in a solvent during a process called solvation. When the solvent is water, the process also is called hydration.
- Every substance has a characteristic solubility in a given solvent.
- Factors that affect solubility include the nature of the solute and solvent, temperature, and pressure.
- Henry's law states that the solubility (S) of a gas in a liquid is directly proportional to the pressure (P) of the gas above the liquid at a given temperature.

15.2 Solution Concentration

- The concentration of a solution is a quantitative measure of the amount of solute in a given amount of solvent or solution.
- Measures of concentration include mass and volume percentages, molarity, molality, and mole fraction.

- A dilute solution can be prepared from a more concentrated standard stock solution.

15.3 Colligative Properties of Solutions

- Physical properties affected by the concentration of the solute but not the nature of the solute are called colligative properties.
- Colligative properties of solutions include vapor pressure lowering, boiling point elevation, freezing point depression, and osmotic pressure.

15.4 Heterogeneous Mixtures

- One of the key differences between solutions, colloids, and suspensions is particle size.
- The random motion of colloidal dispersions due to molecular collisions is called Brownian motion.
- The scattering of light by colloidal particles is called the Tyndall effect. The Tyndall effect can be used to distinguish colloids from solutions.

Key Equations and Relationships

- Henry's law: $\frac{S_1}{P_1} = \frac{S_2}{P_2}$ (p. 460)
- Percent by mass = $\frac{\text{mass of solute}}{\text{mass of solution}} \times 100$ (p. 463)
- Percent by volume = $\frac{\text{volume of solute}}{\text{volume of solution}} \times 100$ (p. 464)
- Molarity (M) = $\frac{\text{moles of solute}}{\text{liters of solution}}$ (p. 464)
- Molarity-volume relationship: $M_1V_1 = M_2V_2$ (p. 467)
- Molality (m) = $\frac{\text{moles of solute}}{\text{kilogram of solvent}}$ (p. 469)
- Mole fractions: $X_A = \frac{n_A}{n_A + n_B}$ $X_B = \frac{n_B}{n_A + n_B}$ (p. 470)
- Boiling point elevation: $\Delta T_b = K_b m$ (p. 472)
- Freezing point depression: $\Delta T_f = K_f m$ (p. 473)

Vocabulary

- | | | |
|--------------------------------------|-------------------------------|------------------------------------|
| • boiling point elevation (p. 472) | • immiscible (p. 454) | • solubility (p. 457) |
| • Brownian motion (p. 478) | • insoluble (p. 454) | • soluble (p. 454) |
| • colligative property (p. 471) | • miscible (p. 454) | • solvation (p. 455) |
| • colloid (p. 477) | • molality (p. 469) | • supersaturated solution (p. 459) |
| • concentration (p. 462) | • molarity (p. 464) | • suspension (p. 476) |
| • freezing point depression (p. 473) | • mole fraction (p. 470) | • Tyndall effect (p. 479) |
| • heat of solution (p. 457) | • osmosis (p. 475) | • unsaturated solution (p. 458) |
| • Henry's law (p. 460) | • osmotic pressure (p. 475) | • vapor pressure lowering (p. 472) |
| | • saturated solution (p. 458) | |

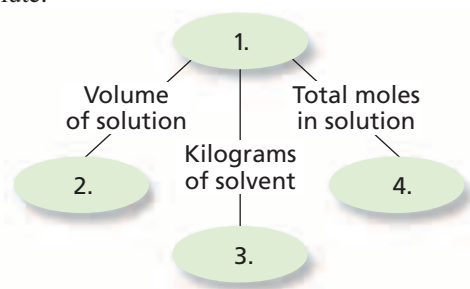




Go to the Chemistry Web site at chemistrymc.com for additional Chapter 15 Assessment.

Concept Mapping

47. Complete the following concept map using the following terms: molarity, mole fraction, molality, moles of solute.



Mastering Concepts

48. What is the difference between solute and solvent? (15.1)
49. What determines whether a solute will be soluble in a given solvent? (15.1)
50. Explain the difference between saturated and unsaturated solutions. (15.1)
51. What does it mean if two liquids are said to be miscible? (15.1)
52. What are three ways to increase the rate of solvation? (15.1)
53. Why are gases less soluble at higher temperatures? (15.1)
54. What is the difference between percent by mass and percent by volume? (15.2)
55. What is the difference between molarity and molality? (15.2)
56. Explain on a particle basis why the vapor pressure of a solution is lower than a pure solvent. (15.3)
57. How does a solute affect the boiling point of a solution? (15.3)
58. How does a solute affect the freezing point of a solution? (15.3)
59. Describe osmosis. (15.4)
60. What is a colligative property? (15.4)

61. What is a suspension and how does it differ from a colloid? (15.4)
62. Name a colloid formed from a gas dispersed in a liquid. (15.4)
63. How can the Tyndall effect be used to distinguish between a colloid and a solution? Why? (15.4)

Mastering Problems

Henry's Law (15.1)

64. The solubility of a gas in water is 0.22 g/L at 20.0 kPa of pressure. What is the solubility when the pressure is increased to 115 kPa?
65. The solubility of a gas in water is 0.66 g/L at 15 kPa of pressure. What is the solubility when the pressure is increased to 40.0 kPa?
66. The solubility of a gas is 2.0 g/L at 50.0 kPa of pressure. How much gas will dissolve in 1 L at a pressure of 10.0 kPa?
67. The solubility of a gas is 4.5 g/L at a pressure of 1.0 atm. At what pressure will there be 45 g of gas in 1.0 L of solution?
68. The partial pressure of CO₂ inside a bottle of soft drink is 4.0 atm at 25°C. The solubility of CO₂ is 0.12 mol/L. When the bottle is opened, the partial pressure drops to 3.0×10^{-4} atm. What is the solubility of CO₂ in the open drink? Express your answer in grams per liter.

Percent Solutions (15.2)

69. Calculate the percent by mass of 3.55 g NaCl dissolved in 88 g water.
70. Calculate the percent by mass of benzene in a solution containing 14.2 g of benzene in 28.0 g of carbon tetrachloride.
71. What is the percent by volume of 25 mL of methanol in 75 mL of water?
72. A solution is made by adding 1.23 mol KCl to 1000.0 g of water. What is the percent by mass of KCl in this solution?
73. What mass of water must be added to 255.0 g NaCl to make a 15.00 percent by mass aqueous solution?
74. The label on a 250-mL stock bottle reads "21.5% alcohol by volume." What volume of alcohol does it contain?
75. A 14.0 percent by mass solution of potassium iodide dissolved in water has a density of 1.208 g/mL. How many grams of KI are in 25.0 mL of the solution?

Molarity (15.2)

- 76.** What is the molarity of the following solutions?
- 2.5 mol KCl in 1.0 L of solution
 - 1.35 mol H_2SO_4 in 245 mL of solution
 - 0.875 mol of ammonia in 155 mL of solution
- 77.** What is the molarity of the following solutions?
- 0.96 g MgCl_2 in 500 mL of solution
 - 9.33 g Na_2S in 450 mL solution
 - 2.48 g CaF_2 in 375 mL of solution
- 78.** How many moles of solute are contained in the following solutions?
- 15.25 mL 2.10M CaCl_2
 - 125 mL 0.0500M $\text{Ba}(\text{OH})_2$
 - 53.1 mL 12.2M HCl
- 79.** How many grams of solute are contained in the following solutions?
- 64.3 mL 0.0238M KOH
 - 142 mL 1.40M K_2SO_4
 - 750.0 mL 0.225M NH_4OH

Molar Dilution (15.2)

- 80.** How many milliliters of 2.55M NaOH is needed to make 125 mL 0.75M NaOH solution?
- 81.** How many milliliters of 0.400M HBr solution can be made from 50.0 mL of 8.00M HBr solution?
- 82.** What is the molarity of each resulting solution when the following mixtures are prepared?
- 500.0 mL H_2O is added to 20.0 mL 6.00M HNO_3
 - 30.0 mL 1.75M HCl is added to 80.0 mL 0.450M HCl

Molality and Mole Fraction (15.2)

- 83.** Calculate the molality of the following solutions.
- 15.7 g NaCl in 100.0 g H_2O
 - 20.0 g CaCl_2 in 700.0 g H_2O
 - 3.76 g NaOH in 0.850 L H_2O
- 84.** Calculate the mole fraction of NaCl, CaCl_2 , and NaOH in the solutions listed in the previous problem.
- 85.** What are the molality and mole fraction of solute in a 35.5 percent by mass aqueous solution of formic acid (HCOOH)?

Colligative Properties (15.3)

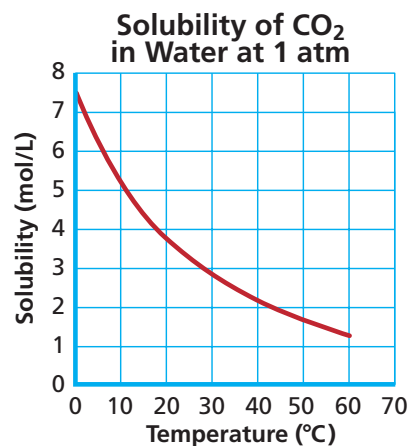
- 86.** Using the information in Tables 15-4 and 15-5, calculate the freezing point and boiling point of 12.0 g of glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) in 50.0 g H_2O .

- 87.** Using the information in Tables 15-4 and 15-5, calculate the freezing point and boiling point of each of the following solutions.
- 2.75m NaOH in water
 - 0.586m of water in ethanol
 - 1.26m of naphthalene (C_{10}H_8) in benzene
- 88.** A rock salt (NaCl), ice, and water mixture is used to cool milk and cream to make homemade ice cream. How many grams of rock salt must be added to water to lower the freezing point 10.0°C ?
- 89.** Calculate the freezing point and boiling point of a solution that contains 55.4 g NaCl and 42.3 g KBr dissolved in 750.3 mL H_2O .

Mixed Review

Sharpen your problem-solving skills by answering the following.

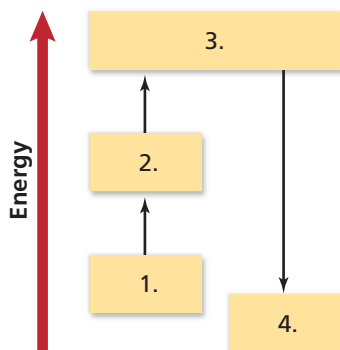
- 90.** If you prepared a saturated aqueous solution of potassium chloride at 25°C and then heated it to 50°C , would you describe the solution as unsaturated, saturated, or supersaturated? Explain.
- 91.** Use the graph below to explain why a carbonated beverage does not go flat as quickly when it contains ice.



- 92.** Which of the following substances will be soluble in the nonpolar solvent carbon tetrachloride (CCl_4): Br_2 , C_6H_{14} , NaNO_3 , HCl? Explain.
- 93.** How many grams of calcium nitrate ($\text{Ca}(\text{NO}_3)_2$) would you need to prepare 3.00 L of a 0.500M solution?
- 94.** What would be the molality of the solution described in the previous problem?

Thinking Critically

- 95. Inferring** Why not spread a nonelectrolyte on a road to help ice melt?
- 96. Using Scientific Diagrams** Complete the diagram below using the following phrases: solution, separated solvent + solute, separated solvent + separated solute, solvent + solute. Is the process described exothermic or endothermic?



- 97. Designing an Experiment** You are given a sample of a solid solute and three aqueous solutions containing that solute. How would you determine which solution is saturated, unsaturated, and supersaturated?
- 98. Using Graphs** The following solubility data was collected in an experiment. Plot a graph of the molarity of KI versus temperature. What is the solubility of KI at 55°C?



Table 15-7

Solubility of KI Data	
Temperature (°C)	Grams of KI per 100.0 g solution
20	144
40	162
60	176
80	192
100	206

- 99. Comparing** Which of the following solutions has the highest concentration? Rank the solutions from the greatest to the smallest boiling point depression. Explain your answer.
- 0.10 mol NaBr in 100.0 mL solution
 - 2.1 mol KOH in 1.00 L solution
 - 1.2 mol KMnO₄ in 3.00 L solution

Writing in Chemistry

- 100.** Investigate the total amount of salt used in the U.S. Construct a circle graph showing the different uses and amounts. Discuss each of these areas in detail. Salt was once used as a currency of high value. Find out why this was the case.
- 101.** Look up the various electrolytes in the human blood stream and discuss the importance of each.
- 102.** Research the contents of the tank scuba divers typically use. How does its composition differ from the air that you breathe? What is the condition known as the bends? How is it treated?

Cumulative Review

Refresh your understanding of previous chapters by answering the following.

- 103.** The radius of an argon atom is 94 pm. Assuming the atom is spherical, what is the volume of an argon atom in nm³? $V = 4/3\pi r^3$ (Chapter 2)
- 104.** Identify which of the following molecules is polar. (Chapter 9)
- SiH₄
 - NO₂
 - H₂S
 - NCl₃
- 105.** Name the following compounds. (Chapter 8)
- NaBr
 - Pb(CH₃COO)₂
 - (NH₄)₂CO₃
- 106.** A 12.0-g sample of an element contains 5.94×10^{22} atoms. What is the unknown element? (Chapter 11)
- 107.** Pure bismuth can be produced by the reaction of bismuth oxide with carbon at high temperatures.
- $$2\text{Bi}_2\text{O}_3 + 3\text{C} \rightarrow 4\text{Bi} + 3\text{CO}_2$$
- How many moles of Bi₂O₃ reacted to produce 12.6 moles of CO₂? (Chapter 12)
- 108.** A gaseous sample occupies 32.4 mL at -23°C and 0.75 atm. What volume will it occupy at STP? (Chapter 14)



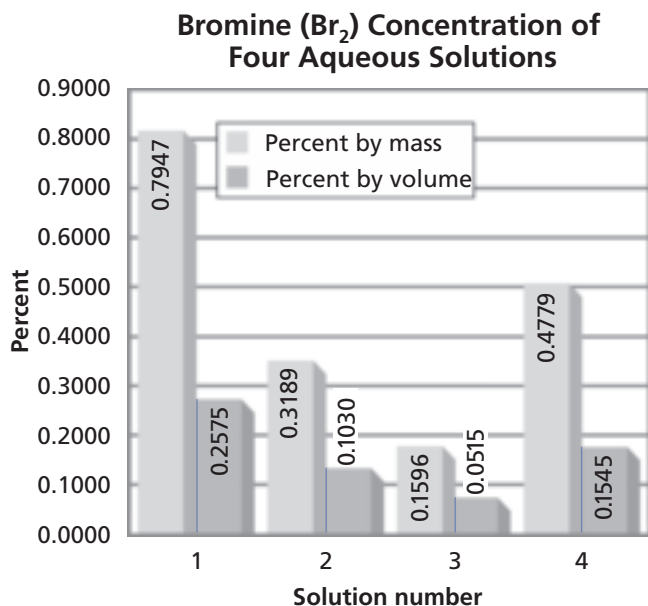
STANDARDIZED TEST PRACTICE

CHAPTER 15

Use these questions and the test-taking tip to prepare for your standardized test.

- How much water must be added to 6.0 mL of a 0.050M stock solution to dilute it to 0.020M?
 - 15 mL
 - 9.0 mL
 - 6.0 mL
 - 2.4 mL
- At a pressure of 1.00 atm and a temperature of 20°C, 1.72 g CO₂ will dissolve in 1 L of water. How much CO₂ will dissolve if the pressure is raised to 1.35 atm and the temperature stays the same?
 - 2.32 g/L
 - 1.27 g/L
 - 0.785 g/L
 - 0.431 g/L
- What is the molality of a solution containing 0.25 g of dichlorobenzene (C₆H₄Cl₂) dissolved in 10.0 g of cyclohexane (C₆H₁₂)?
 - 0.17 mol/kg
 - 0.014 mol/kg
 - 0.025 mol/kg
 - 0.00017 mol/kg
- If 1 mole of each of the solutes listed below is dissolved in 1 L of water, which solute will have the greatest effect on the vapor pressure of its respective solution?
 - KBr
 - C₆H₁₂O₆
 - MgCl₂
 - CaSO₄
- What volume of a 0.125M NiCl₂ solution contains 3.25 g NiCl₂?
 - 406 mL
 - 201 mL
 - 38.5 mL
 - 26.0 mL

Interpreting Graphs Use the graph to answer questions 6–8.



- What is the volume of bromine (Br₂) in 7.000 L of Solution 1?
 - 55.63 mL
 - 8.808 mL
 - 18.03 mL
 - 27.18 mL
- How many grams of Br₂ are in 55.00 g of Solution 4?
 - 3.560 g
 - 0.084 98 g
 - 1.151 g
 - 0.2628 g
- Which of the following relationships is true?
 - 2 × Concentration of solution 2 = Concentration of solution 3
 - 0.5 × Concentration of solution 2 = Concentration of solution 3
 - Concentration of solution 2 = 0.25 × Concentration of solution 3
 - Concentration of solution 2 = 3 × Concentration of solution 3
- Which is NOT a colligative property?
 - boiling point elevation
 - freezing point depression
 - vapor pressure increase
 - osmotic pressure
- How can colloids be distinguished from solutions?
 - dilute colloids have particles that can be seen with the naked eye
 - colloid particles are much smaller than solvated particles
 - colloid particles that are dispersed will settle out of the mixture in time
 - colloids will scatter light beams that are shone through them

TEST-TAKING TIP

Take a Break! If you have a chance to take a break or get up from your desk during a test, take it! Getting up and moving around will give you extra energy and help you clear your mind. During your stretch break, think about something other than the test so you'll be able to get back to the test with a fresh start.

